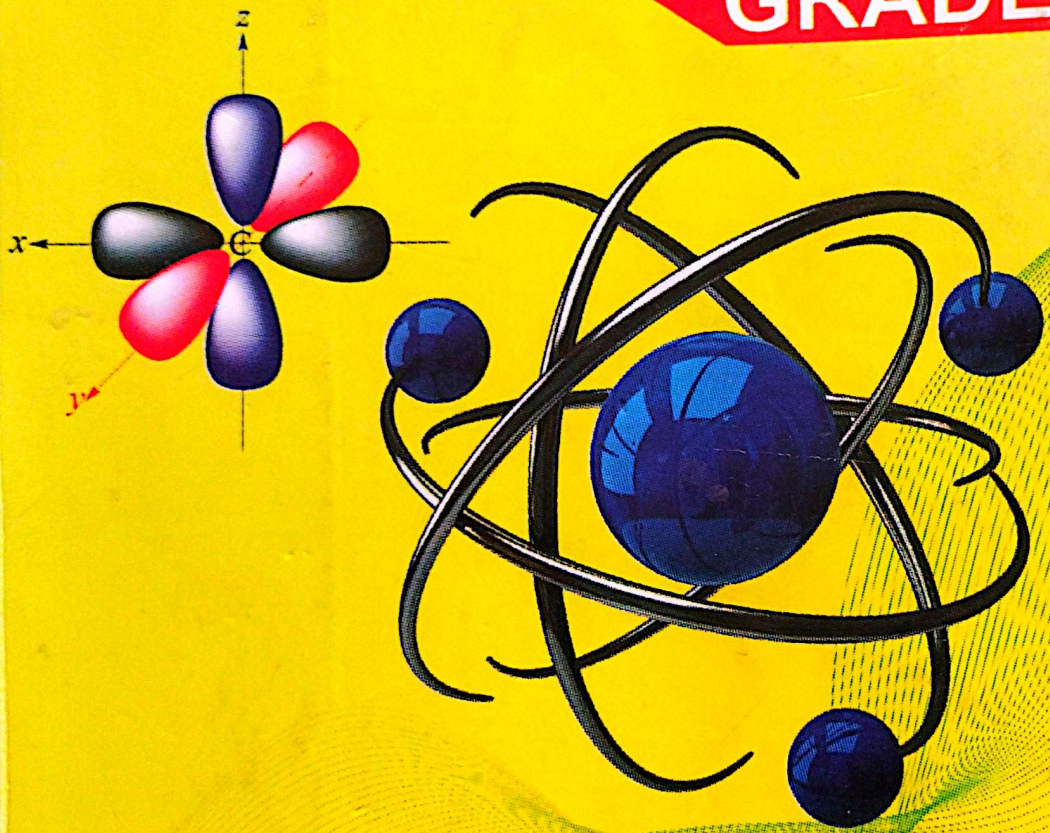
Extreme Series

CHEMISTRY

GRADE 9 - 10



Based on Common Currently Used
National and International Curriculum

Extreme Series

2017 E.C

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UNIT 1

Chemistry and Its Importance

Introduction

Chemistry is the science of matter and its transformation. Chemistry is important and interesting subject that touches our lives everyday.

1.1 Definition and Scope of Chemistry

Definition of Chemistry

Chemistry is a branch of natural science which deals with properties, composition, structure and transformation of substances.

- ☑ **Property** means behavior or characteristics of substances such as color, odor, state, reactivity and so on.
- ☑ **Composition** means the particles that make up a given substance. For example, water is composed of hydrogen and oxygen.
- ☑ **Structure** of a substance means how the components of the substance are arranged in it.

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☑ **Transformation** is the changes that the substance undergoes.

☑ **Substance** is a form matter that has a definite or constant composition and distinct properties.

☑ **Matter** is anything that occupies space and possesses mass.

Example:

- water
- air
- salt
- stone etc

The scope of chemistry is the extent of the application of chemistry in discovering and understanding of matter to improve the living condition for mankind.

The importance of chemistry can be summarized in short by the figure below.

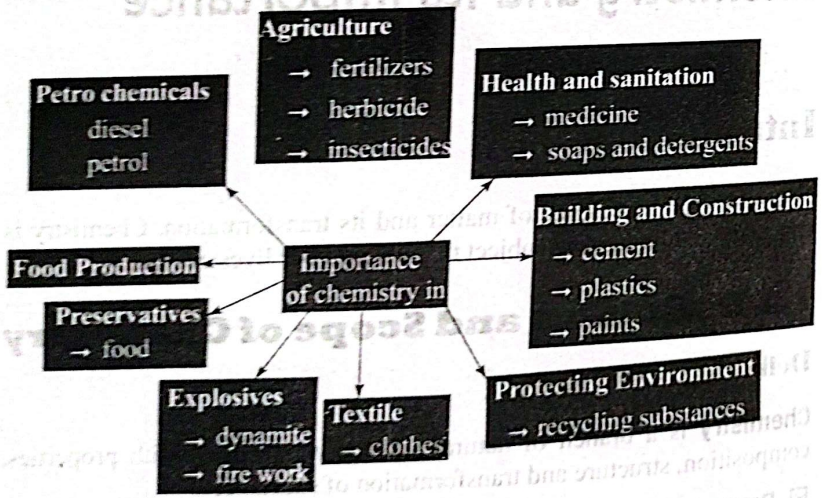

Fig 1 .1 Importance of Chemistry

Illustrative Examples

1. Sugar is sweet. This description tells us _____ of sugar.

A. composition C. structure

B. property D. transformation

Solution:

Color, odour, taste, reactivity and state are terms used to describe property of substance.

Answer: B

2. The ingredients or constituents in table salt are sodium and chlorine. What does this indicate?

A. property of table salt C. structure of table salt

B. transformation of table salt D. composition of table salt

Solution:

Composition of a substance represents its ingredients or constitutents.

Answer: D

3. The sanitizer we use for covid-19 belongs to:

A. Antibiotic C. Disinfectants

B. Tranquillizers D. Aesthetics

Answer: C

4. Explain the importance of chemistry in :

a. Agriculture c. Construction

b. Food preservation d. Medicine

Solution:

a. To increase crop yield, chemistry provides through chemical industries, the following important chemicals and materials.

- Fertilizers
- Herbicides
- Insecticides
- Plastic pipes for irrigation

b. Chemistry helps to preserve food for a long period of time.

c. Chemistry supplies construction or building chemicals and materials for bridges, roads, house and factory buildings. The chemicals can be cement, paint and glues while the materials may be plastics and steels.

d. Chemistry is important to manufacture different medicines, such as:

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- **Antibiotics** - control infection and diseases
- **Analgesics** - pain killers
- **Antiseptics** - stop contamination of wound.
- **Disinfectants** - kill microbe in floors, drain and toilets.
- **Tranquillizers** - reduce tension to patients suffering from mental diseases.
- **Anaesthetics** - give relief during medical operation to the patient.

5. Write some examples of non-matter things

Solution:

- Sound, love, thoughts, time, idea and sunlight are non-matter things that do not have mass and occupy space.

6. Chemistry deals with the _____, _____, _____ and _____ of matter.

Solution:

Composition, structure, property and transformation

7. Explain how chemistry touches our lives.

Solution:

Chemistry touches almost every aspects of our lives, because the air we breath, the food we eat, the fluid we take, our clothes, our transportation and fuel supplies need the knowledge of chemistry directly or indirectly.

1.2 Major Branches of Chemistry

Chemistry has five major branches. These are:

1. **Inorganic chemistry** is the study of all elements and inorganic compounds.
2. **Organic chemistry** deals with all compounds of carbon excepts oxides of carbon (CO and CO_2), carbonates (CO_3^{2-}), hydrogen carbonates (HCO_3^-), cyanides (CN^-), cyanates (CNO^-) and thiocyanates (CNS^-)

3. **Analytical chemistry** is concerned with quantitative determination and qualitative identification of a sample of matter.
4. **Physical chemistry** is the study of macroscopic and microscopic properties of a chemical system.

Macroscopic properties - large scale properties without explicit use of molecules.

Microscopic properties - small scale properties with explicit use of molecules.

- → Thermodynamics
- → Spectroscopy
- → Quantum chemisrty
- → Chemical kinetics
- → Electro chemistry

} are sub – branches of physical chemistry

5. **Biochemistry** bridges chemistry and biology.

☑ It is the study of the science of the molecules and chemical reactions in living things.

☑ It utilizes the principles and language of chemistry to explain biology at molecular level.

The Relationship Between Chemistry and other Natural Sciences

Chemistry overlaps with the common natural sciences as follows:

- Chemistry + Physics = Physical chemistry
- Chemistry + Biology = Biochemistry
- Chemistry + Geology = Geo chemistry
- Chemistry + Medicine = medicinal chemistry

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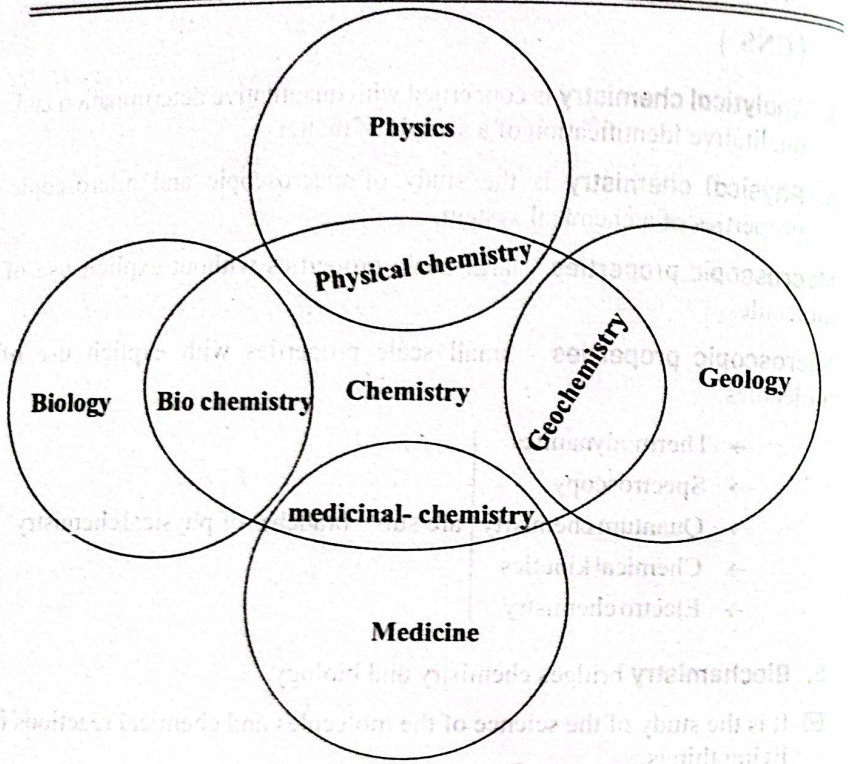

Fig 1 .2 The Relationship Between Chemistry and other Natural Sciences

Illustrative Examples

8. A chemistry laboratory technician identified a substance that polluted a water from a sample of it and determined the mass of the substance. Which of the following chemistry is very important for this practical work?

A. Physical chemistry C. Organic chemistry

B. Analytic chemistry D. Inorganic chemistry

Answer: B

9. Thermodynamics, spectroscopy and electrochemistry are parts of:

A. analytical chemistry B. bio chemistry

C. organic chemistry D. physical chemistry

Answer: D

10. Which of the following substances affect human being's life?

A. chlorofluorocarbons C. antibiotics

B. fertilizers D. cement

Solution:

Chemical industries produce useful and harmful substances. Some of the dangerous substances that affect human's life are chlorofluorocarbons, oxides of carbon, nitrogen and sulphur.

Answer: A

11. All of the following are agricultural chemicals. **EXCEPT**

A. Fertilizers C. Paints

B. Herbicides D. Insecticides

Solution:

Paints are chemicals used in construction.

Answer: C

12. Why chemistry is called "central science"?

Solution:

Because it is linked with other natural sciences.

1.3 Some Common Chemical Industries in Ethiopia

Industry is an institution that converts raw materials into semi-finished and finished products.

- ☑ The starting materials (raw materials) are called **reactants** and the new substances formed are known as **products**.
- ☑ An industry may contain a group of companies that produce different products. Such group of companies are called **industrial parks**. There are many industrial parks in our country, Ethiopia. Some of them are listed below

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- Hawassa industrial park
- Bole Lemi industrial park
- Kombolcha industrial park
- Dire Dawa industrial park
- Mekele industrial park
- Debre Birhan industrial park

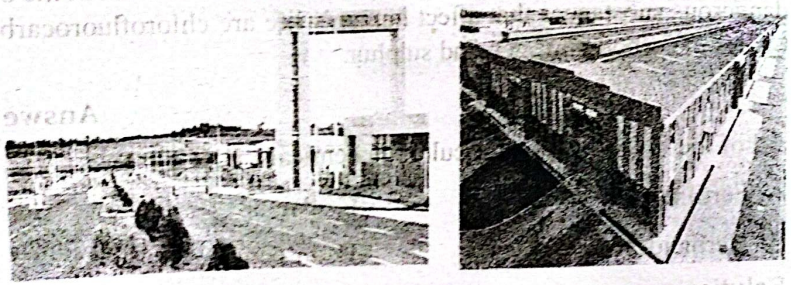

Fig 1 .3 Chemical Industry

A chemical industry is involved in the manufacturing of chemical products such as:

- fertilizers
- agro chemicals
- pharmaceuticals
- medical chemicals
- plastics
- rubber products
- botanical products
- cement
- paints
- printing inks
- detergents
- perfumes
- soaps

Table : Some Chemical Industries Found in Ethiopia

No	Industry	Location	Products
1	Chora gas factory	Addis Ababa	Oxygen, acetylene, floor wax, shoe polish
2	Nefas silk paints	Addis Ababa	Paints
3	Plastic factory	Addis Ababa	Plastics
4	Glass factory	Addis Ababa	Glasses
5	Foam factory	Addis Ababa	Foam
6	Tannery	Addis Ababa	leather products
7	Moha soft drinks	Addis Ababa	Soft drinks (Pepsi, miranda, 7-up)
8	Kadisco chemical industry	Addis Ababa	Paint
9	Y.B cosmetics	Addis Ababa	Cosmetics
10	Mekabe PLC (cosmetics)	Addis Ababa	Cosmetics
11	Petera soap factory	Oromia (Alemgena)	Soap
12	Teamco hi-jaxs soap factory	Oromia (Barayu)	Soap
13	Zeway caustic soda	Oromia (Adamitulu)	Caustic soda
14	Awash Melksa Aluminium sulphate factory	Oromia (AwashMelkasa)	Sulphuric acid and aluminium sulphate
15	Adola Magnesium oxide factory	Oromia (Adolarede)	Magnesium oxide
16	Abijata soda Ash factory	Oromia (Bulbula)	Soda ash
17	Adami-Tulu pesticide factory	Oromia (AdamiTulu)	Pesticides
18	Zemenawi Building industries	Oromia (Dukem)	Building materials

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19	Mesfin Tadesse soap factory	Oromia (Woliso)	Soap
20	Get Eshet Detergent	Oromia (Bishoftu)	Detergents
21	Natural Gum Manceting Enterprise	Amhara (Awi)	Gum
22	Fanos detergent soap factory	Tigray (Axum)	Detergents
23	Etab laundry soap factory	Sidama (Hawassa)	Laundry soap
24	Cement factory	Mugher, Dire Dawa, Mesobo, Derba, Dangote	Cement
25	Paper and pulp	Wonji	Paper and pulp
26	Sugar	Wonji, Metehara, Fincha, Omokuraz	Sugar, ethanol
27	Epharm	Addis Ababa, Adigrat	Pharmaceuticals
28	Tyre factory	Addis Ababa	Tyres
29	Textile factories	Kombolcha, Hawassa, Bahir Dar, Arba mich, Dire Dawa, Adwa	Textiles

Illustrative Examples

13. Which of the following chemical industries in Ethiopia mainly serves as for cleaning purpose?

A. Cement C. Detergent

B. Paint D. Agro chemical

Solution:

Cement and paint - building materials

Agro chemicals are important to increase crop yields.

Answer: C

14. Aluminum sulfate factory is found in

A. Addis Ababa C. Wonji

B. Hawassa D. Awash Melkasa

Answer: D

15. Which of the following is not a chemical industry?

A. cement factory C. biscuit factory

B. soap and detergent factory D. soda ash factory

Solution:

- Chemical factories produce chemical products such as cement, paint, soda ash, soap.
- Food, oil and sugar factories produce non-chemical products.

Answer: C

16. Which of the following chemical industry is found in Addis Ababa?

A. Paint factory C. Soda ash factory

B. Cement factory D. Sulphuric acid factory

Answer: A

17. List some of the products of companies in a chemical industry.

Solution:

Inorganic and organic substances such as soaps and detergents, explosives, agrochemicals, polymers, ceramic products, petro chemicals and polymers.

18. What are oileochemicals?

Solution:

Oils, fats and waxes collectively known as oileo chemicals.

19. Write some natural raw materials used by chemical industries.

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Solution:

Natural gas, water, air, metals, minerals, oil, animals skin, plants.

20. What are chemical products?

Solution:

Chemical products are products produced, processed and distributed by the company.

21. Chemistry deals with composition of substances. Identify the components of table sugar that we use in our drink.

Solution:

Table sugar is chemically known as sucrose contains carbon, hydrogen and oxygen. Its molecular formula is $C_{12}H_{22}O_{11}$.

22. Explain the difference between substance and matter.

Solution:

A substance is any kind of matter possessing constant physical properties under specific conditions. However, matter is any thing that occupies space and has mass.

23. Water is colorless, odorless and tasteless liquid in its pure form at room temperature. This description tells us _____ of water.

Solution:

Property of water

24. List some non-chemical industries found in Ethiopia.

Solution:

- Garment industry
- Textile factory
- Flour industry
- Food factory
- Vegetable factory

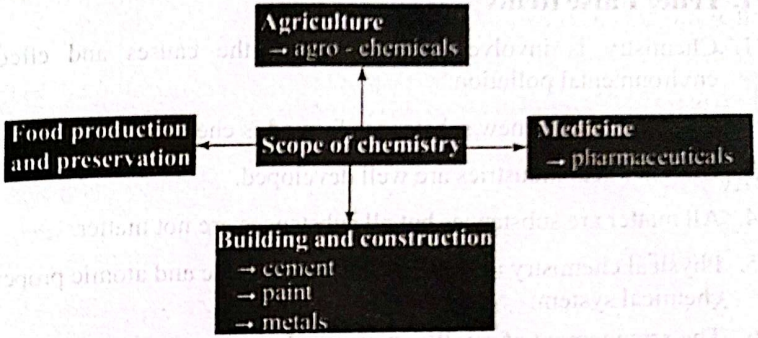
25. How chemists help farmers?

Solution:

Chemists help farmers by producing fertilizers, insecticides and herbicides that increase crop yields.

26. The scope of chemistry includes four major groups. Explain by using diagram.

Solution:



27. Write some branches natural science.

Solution:

- Chemistry
- Biology
- Astronomy
- Physics
- Geology
- Earth science

28. Chemistry is a science that touches our life every moment. Explain how?

Solution:

Because chemistry has a role in obtaining useful substances such as drugs, fertilizers, soaps, textiles, cosmetics, plastics and beverages directly or indirectly linked to our life.

29. How chemistry helps the construction industry?

Solution:

By manufacturing cement, limestone, iron, aluminium, glass, plastics and paints.

30. Why chemists recycle used can?

Solution:

- To protect the environment from pollution.
- For economical utilization of natural resources.

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Supplementary Exercise

I. True/ False items

1. Chemistry is involved in discovering the causes and effects of environmental pollution.
2. Reactants are the new substances formed is chemical industry.
3. Our chemical industries are well developed.
4. All matter are substances but all substances are not matter.
5. Physical chemistry is the study of macroscopic and atomic properties a chemical system.
6. The arrangement of small particles such as atoms or ions in a substance is its structure.
7. Soaps and detergents are dirt removers while insecticides are weed killers.
8. Chemistry is important to fulfill our basic needs such as food, cloth and shelter.
9. Medicinal chemistry is concerned to make pharmaceuticals and drugs.
10. Radiation, light, sound and knowledge are examples of matter.

II. Multiple Choice Items.

11. The subject that deals with structure of proteins and associates with energy is:

A. Physical chemistry C. Physical chemistry

B. Water chemistry D. Biochemistry

12. Which of the following is a plant nutrient?

A. Fertilizer B. Insecticide C. Herbicide D. Pesticide

13. The product of chemical industry that can be used as pain killer is

A. Penicillin B. ORS C. Paracetamol D. Aesthetic

14. Which of the following is not the importance of chemistry?

A. Protecting and monitoring the environment

B. Producing toxic substances that deplete the ozone layer

C. Manufacturing lighting, heating and photography substances.

D. Treating drinking water with chlorine.

15. All of the following are basic chemical products. **EXCEPT.**

A. acids and bases C. salts

B. organic chemicals D. plastic materials

16. Which of the following kills microbes present in toilets, floors and drains?

A. Antibiotics C. Disinfectants

B. Antisepses D. Analgesics

17. Which of the following is not a pest that damage crops?

A. bee B. aphid C. whitefly D. caterpillar

18. Which of the following is irreversible transformation in your home?

A. Freezing of water in refrigerator.

B. Dissolution of sugar in a cup of tea.

C. Cooking of rice for breakfast.

D. Boiling of water for coffee.

19. The _____ of a substance is its attribute, quality of character.

A. transformation C. structure

B. property D. composition

20. All of the following substances are dangerous substances that can affect human's life and the environment. **EXCEPT.**

A. Chlorofluorocarbons C. Carbon monoxide

B. Oxides of nitrogen D. Pharmaceuticals

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III. Matching items

A	B
21. Analytical chemistry	A. composition of minerals in earth
22. Biochemistry	B. synthesis of pharmaceuticals
23. Geochemistry	C. area of research and modelling
24. Medicinal chemistry	D. metabolic processes
25. Mathematical chemistry	E. qualitative identification and quantitative determination
	F. hydrocarbons and their derivatives.

IV. Short Answer Items

26. Write four major field that need the knowledge of chemistry.
27. List some sub-branches of physical of chemical products?
28. The study of chemistry involves _____ and _____ informations.
29. A form of matter that has constant compositions is _____ .
30. What are the two types of transformations?

Answer for Supplementary Exercise

Part I

1. True	3. False	5. True	7. False	9. True
2. False	4. False	6. True	8. True	10. False

Part II

11. D	13. C	15. D	17. A	19. B
12. A	14. B	16. C	18. C	20. D

Part III

21. E	22. D	23. A	24. B	25. C

Part IV

25. Medicine
    - ➥ Food production
    - ➥ Agriculture
    - ➥ Building and construction

26. Chemical kinetics
    - ➥ Spectroscope
    - ➥ Thermodynamics
    - ➥ Electrochemistry
    - ➥ Quantum chemistry

27. Basic chemicals
    - ➥ Chemical products to be used for further manufacture.
    - ➥ Finished chemicals to be used for ultimate consumption.

28. Substance

29. Chemical and physical transformations

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UNIT 2

Measurement and Scientific Method

Introduction

Some of our observation of physical properties of substances are qualitative or descriptive. However, advances in science often depend on quantitative observations, which depend on measurements and reporting with numerical values.

2.1 Measurement

- It is the act of finding the magnitude of some quantity interms of a previously defined unit that serves as a standard.
- A measured quantity has magnitude and unit tied to it.

Example:

The mass of an object is 10kg, 10 is magnitude, kg is unit.

- Measurement of properties may be referred as macroscopic property or microscopic property.
- **A macroscopic property** is a property that can be determined directly

Example:

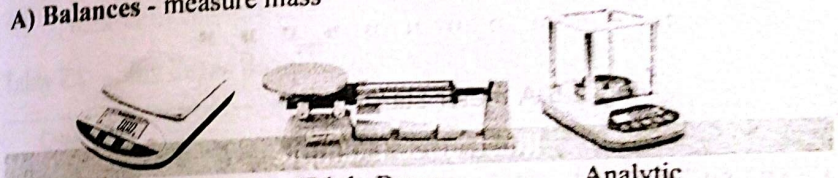
Measurement of mass, volume, thickness of an object.

A microscopic property is a property of substance that can be determined indirectly.

Example: Charge of an electron

The common devices we use in chemistry laboratory to make simple measurements are:

A) Balances - measure mass



Electronic Balance | Triple Beam Balance | Analytic Balance

Fig 2.1 Different Balances

B) pH meter- measures pH of solution

Fig 2.2 pH Meter

C) Burette, pipette, measuring cylinder and volumetric flask- used to measure volume of liquids in laboratory

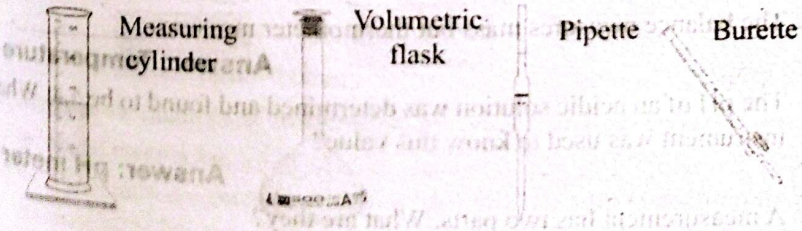


Fig 2.3 Commonly used Devices in Chemistry Laboratory

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to measure volume of liquids

D) Thermometer - is used to measure temperature

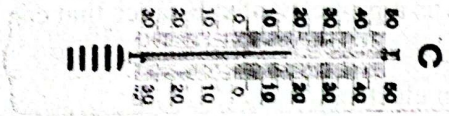

Fig 2.4 A thermometer

E) Meter stick - to measure the length of an object.

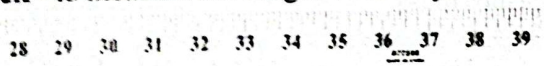

Fig 2.5 A meter stick

Illustrative Example

1. All of the following devices are used to measure volume of liquids in chemistry laboratory. **EXCEPT**

   A) Pipette C) Volumetric flask

   B) Burette D) Beam balance

   Solution:

   Beam balance is used to measure the mass of a sample of matter.

   Answer: D

2. Traditionally, the length of a room can be measured with

   A) Arm B) Foot C) Hand span D) All

   Solution:

   Traditionally, length can be measured using, arm, foot, cubit, hand span and so on.

   Answer: D

3. The balance measures mass but thermometer measures _____.

   Answer: Temperature

4. The pH of an acidic solution was determined and found to be 5.4. What instrument was used to know this value?

   Answer: pH meter

5. A measurement has two parts. What are they?

Solution:

Magnitude and unit.

International System of Units (SI Units)

- Measurements in the scientific world are usually expressed in **SI units.**
- The SI unit is based on the **seven fundamental or basic units**
- **A basic unit** is one that is defined by a physical standard of measurement.

Table 2.1 The Seven Basic Quantities and their SI Units

Quantity	Name of SI unit	Symbol of SI unit
Length	Meter	m
Mass	Kilogram	kg
Time	Second	s
Temperature	Kelvin	K
Electric current	Ampere	A
Amount of substance	Mole	mol
Luminous intensity	Candela	Cd

i. **Length** is the measurement or extent of something from one end to another end.

- Its SI unit is meter (m)
- Other units of length can be **kilometer (km), decimeter (dm), centimeter(cm), millimeter(mm)** and so on.

1km = 1000m = 10,000dm = 100,000cm = 1,000,000mm

1m = 10dm = 100cm = 10,000mm

1dm = 10cm = 100mm

1cm = 10mm

ii. **Mass** is the quantity of matter in an object.

- Its SI units is **kg.**
- Other units are gram(g), decigram(dg), milligram(mg), centigram (cg)

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1kg = 1000g

1g = 10dg = 100cg = 1000mg

1dg = 10cg = 100mg

1cg = 10mg

iii. Temperature is the measure of hotness or coldness of an object.

- It is the measure of the average kinetic energy of molecules.
- Its SI unit is **Kelvin**. Other units of temperature are degree Celsius (°C) and degree Fahrenheit (°F)
- The lowest attainable temperature is **- 273.15°C or 0K**

Relationship Between K and °C.

Subtract the lowest value from both sides

$K - O = {}^\circ C - (-273.15)$

$K = {}^\circ C + 273.15$

For calculation, omit .15.

$K = {}^\circ C + 273$

$^\circ C = K - 273$

Example: convert

a) 37°C to K

b) 298K to °C

Solution:

a) $K = {}^\circ C + 273$

$= 37 + 273$

$= 310K$

b) $^\circ C = K - 273$

$= 298 - 273$

$= 25^\circ C$

Relationship Between °C and °F

Use the boiling and melting points of water to get a formula that relates °C and °F.

- Boiling point of water **100°C or 212°F.**
- Melting or freezing point of water is **0°C or 32°F.**
- Subtract freezing point from both sides:

$$\frac{^\circ C - 0}{^\circ F - 32} = \frac{100 - 0}{212 - 32}$$

$$\Rightarrow \frac{^\circ C}{^\circ F - 32} = \frac{100}{180} = \frac{5}{9}$$

$$\therefore \boxed{^\circ C = \frac{5}{9}(^\circ F - 32)} \qquad \boxed{^\circ F = \left(\frac{9}{5} \times {}^\circ C\right) + 32}$$

Example: convert:

a) 37°C to °F

b) 59°F to °C

Solution:

a) $^\circ F = \frac{9}{5} \times {}^\circ C + 32$

$= \frac{9}{5} \times 37 + 32$

$= 98.6^\circ F$

$\therefore 37^\circ C = 98.6^\circ F$

b) $^\circ C = \frac{9}{5}(^\circ F - 32)$

$= \frac{9}{5}(59 - 32)$

$= 15^\circ C$

$\therefore 59^\circ F = 15^\circ C$

Heat is the transfer of thermal energy between two bodies that are at different temperatures.

Heat flows spontaneously from hot body to cold body.

iv. A mole is the amount of substance that contains as many elementary particles as there are atoms in 12g of carbon-12 isotope.

1mole = 6.02×10^{23} particles (atoms, molecules, ions, formula units)

$$\text{Number of mole} = \frac{\text{given mass}}{\text{molar mass}}$$

$$\boxed{n = \frac{m}{M}}$$

- Number of atoms of element $= n \times 6.02 \times 10^{23} \frac{\text{atoms}}{\text{mol}}$

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- Number of atoms in a molecule

$$= n \times \text{sum of subscripts} \times 6.02 \times 10^{23} \frac{\text{atoms}}{\text{mol}}$$

- Number of molecules $= n \times 6.02 \times 10^{23} \frac{\text{molecules}}{\text{mol}}$

Example: calculate the number mole and atoms in 22g CO_2. (Atomic mass: C = 12, O = 16)

Solution:

$$m = 22g$$

$$M = 1 \times 12 + 2 \times 16 = 44g/mol$$

$$n = \frac{m}{M} = \frac{22g}{44g/mol} = 0.5mol$$

Number of atoms in a compound $= n \times \text{sumsubscript} \times 6.02 \times 10^{23} \frac{\text{atoms}}{\text{mol}}$

$$= 0.5mol \times (1+2) \times 6.02 \times 10^{23} \frac{\text{atoms}}{\text{mol}}$$

$$= 9.03 \times 10^{23} \text{ atoms.}$$

Illustrative Examples

6. Which of the following is not basic SI unit?
   A) meter B) Kelvin C) second D) Gram

Solution:

Gram is not SI unit of mass. Kilogram is its SI unit.

Answer: D

7. Which one of the following is not basic quantity?
   A) Volume C) Time
   B) Temperature D) Current

Answer: A

8. At what numerical value °C is equal to °F?

Solution:

$$°C = \frac{5}{9}(°F - 32)$$

$$°C = \frac{5}{9}(°C - 32) \text{ beacuse } °C = °F$$

$$\frac{9}{5} \cdot °C = -32$$

$$\frac{9}{5}°C - °C = -32$$

$$\frac{4}{5}°C = -32$$

$$°C = -32 \times \frac{5}{4} = -40$$

$$\therefore -40°C = -40°F$$

9. Determine the number of moles of hydrogen atoms in 0.5mol $(NH_4)_2HPO_4$.

Solution:

Number of mole of an element in a compound = mole of compund × subscript of the element

$\Rightarrow$ mole of hydrogen = mole of $(NH_4)_2HPO_4$ × subscript of H

$= 0.5mol \times (2 \times 4 + 1) = 0.5mol \times 9 = 4.5mol$

10. Convert $-40°F$ to Kelvin.

Solution:

$$-40°F = -40°C$$

$$K = °C + 273$$

$$= -40 + 273 = 233K$$

11. Convert 0.685 kg into gram.

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Solution:

$1kg = 1000g$

$0{\cdot}685kg = ?$

$$? = \frac{0{\cdot}685kg \times 1000g}{1kg} = 685g$$

$\therefore 0{\cdot}685kg = 685g$

12. Explain the difference between temperature and heat.

Solution:

Temperature measures the intensity of heat but heat is a form of energy that flows spontaneously from hot body to cold body.

Common prefixes used in SI units

Prefixes are very important in converting one unit into another.

Table 2.2 Common Prefixes Used in SI Units

Prefix	Symbol	Decimal/ Factor
Exa	E	$10^{18} = 1,000,000,000,000,000,000$
Peta	P	$10^{15} = 1,000,000,000,000,000$
Tera	T	$10^{12} = 1,000,000,000,000$
Giga	G	$10^{9} = 1,000,000,000$
Mega	M	$10^{6} = 1,000,000$
Kilo	k	$10^{3} = 1,000$
Hecto	h	$10^{2} = 100$
Deca	da	$10^{1} = 10$
Deci	d	$10^{-1} = 0.1$
Centi	c	$10^{-2} = 0.01$
Milli	m	$10^{-3} = 0.001$
Micro	μ	$10^{-6} = 0.000,001$
Nano	n	$10^{-9} = 0.000,000,001$
Pico	p	$10^{-12} = 0.000,000,000,001$
Femto	F	$10^{-15} = 0.000,000,000,000,001$
Atto	A	$10^{-18} = 0.000,000,000,000,000,001$

Example: $1Ms = 1 \times 10^{6}s$

$1nm = 1 \times 10^{-9}m$

$1mg = 1 \times 10^{-3}g$

Conversion of Units

To convert one unit into another, first find the conversion factor using ratio method.

Example: convert

a) 0.25dm into mm

b) 125nm into cm

Solution:

a) $\frac{1dm}{1mm} = \frac{10^{-1}m}{10^{-3}m}$

$\frac{1dm}{1mm} = 10^{2}$ conversion factor

$\therefore \frac{1dm}{1mm} = 1 \times 10^{2}\frac{mm}{dm}$

Thus, the conversion factor is 10^2

$\Rightarrow 0{\cdot}25dm = 0{\cdot}25d\not{m} \times 10^{2}\frac{mm}{d\not{m}} = 25mm$

b) $\frac{1nm}{1cm} = \frac{10^{-9}m}{10^{-2}m}$

$\frac{1nm}{1cm} = 10^{-7}$

$\Rightarrow 1nm = 10^{-7}\frac{cm}{nm}$

$\Rightarrow 125nm = 125n\not{m} \times 10^{-7}\frac{cm}{n\not{m}} = 1{\cdot}25 \times 10^{-5}cm$

Illustrative Examples

13. How many milligrams are there in 0.005 hectogram?

A) 50 B) 500 C) 5 D) 0.5

Solution:

$$1mg = 10^{-3}g \qquad \frac{1hg}{1mg} = \frac{10^{2}g}{10^{-3}g}$$

$$1hg = 10^{2}g \qquad \frac{1hg}{1mg} = 10^{5}$$

$$1hg = 10^{5}\frac{mg}{hg}$$

$$\Rightarrow 0\cdot005hg = 0\cdot005\cancel{hg} \times 10^{5}\frac{mg}{\cancel{hg}} = 500mg$$

Answer: B

14. From the seven basic SI units, which one has a prefix?

A) meter B) Kelvin C) kilogram D) second

Solution:

Kilogram is the basic SI unit with a prefix kilo (10^3).

Answer: C

15. Covert $200cm^3$ into dm^3

Solution:

$$\frac{1cm^3}{1dm^3} = \frac{(10^{-2}m)^3}{(10^{-1}m)^3} = \frac{10^{-6}m^3}{10^{-3}m^3} = 10^{-3}$$

$$\Rightarrow 1cm^3 = 10^{-3}\frac{dm^3}{cm^3}$$

$$\therefore 200cm^3 = 200\cancel{cm^3} \times 10^{-3}\frac{dm^3}{\cancel{cm^3}} = 0\cdot200dm^3$$

16. Convert 0.82ng to pg.

Solution:

$$\frac{1ng}{1pg} = \frac{10^{-9}g}{10^{-12}g} = 10^{3}$$

$$\Rightarrow 1ng = \frac{10^{3}pg}{ng}$$

$$0\cdot82ng = 0\cdot82\cancel{ng} \times 10^{3}\frac{pg}{\cancel{ng}} = 820pg$$

17. What fraction of a second is a nano second?

Solution:

10^{-9} second is a nano second.

18. Write the complete meaning of the following abbreviations that are written using prefixes.

a) KW b) km c) MW d) kV

Solution:

a) kilo watts
b) kilo meter
c) Mega watt
d) kilovolt

2.2 Uncertainty in Measurements

In every measurement, there is always some uncertainty in it due to:

- Observational error
- Instrumental error
- Variability in the item measured
- environment where the measurement is made.

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Uncertainity

Systematic

- Causes the value to be too small or too large
- Can be eliminated
- Can be caused by faulty of instruments and improper use of instruments

Random

- Exists without a predictable pattern
- Can be reduced but not eliminated
- Caused by indidual difference between participants, poorely controlled experimental producers.

Calculating Uncertainties

- Many techniques are applied to produce an estimate of the uncertainty in the value of the average or mean.
- Consider the example below: A student measures the length of a table three times in millimeter (mm) with graduated meter stick and obtained the following data.

Trial 1	Trail 2	Trial 3	Trial 4
99mm	100mm	104mm	101mm

The uncertainty in the mean value can be determined by using

i. The Half Range

$$\text{Uncertainity in the mean length} = \frac{104-99}{2} = 2\cdot 5\text{mm}$$

Thus, the length of the table is 101 ± 2.5mm and the uncertainty is reported as ± mm

ii. The Reading Furthest from the Mean

Uncertainty in the mean length 101 – 104 = –3mm

iii. Using the Resolution of the Instrument

This method is applied in two cases:

- For single reading

- For repeated reading that have the same value.
- If the three trials above were all 101mm, the length lies between 100.5mm because the meterstick reads to half a millimeter. Here the uncertainty in length is 0.5mm. Thus, the uncertainty in length is 101mm ± 0.5mm.
- For example, consider a measuring cylinder, we may probably take the reading as 50.0, but in doing so, we say that it is near to 50.0 than it is 49.9 or 50.1. In other words, it is greater than 49.95 and smaller than 50.05. Hence, we should recode this value a as 50.0 ± 0.05. Thus, 0.05 is absolute uncertainty. The uncertainty in a measurement can be expressed either in absolute uncertainty of the last digit or as percentage uncertainty.

$$\%\,\text{uncertainity} = \frac{\text{absolute uncertainity}}{\text{measurement}} \times 100$$

Therefore, absolute uncertainty = 50.0 ± 0.05

percentage uncertainty = 50.0 ± 0.1%

Illustrative Example

19. What are the causes for uncertainty in measurements?

Solution:

- The person that measures.
- The instrument used to measure.
- The item to be measured.
- The environment where the measurement is made.

20. What are the two classes of uncertainties?

Solution:

Systematic uncertainty

Random uncertainty

21. A thermometer reads 37.5°C. The absolute uncertainty is 0.2°C. How could you report the result in terms of absolute and percentage uncertainty?

Solution:

- In absolute uncertainty: 37.5°C ± 0.2°C

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$$\%\,\text{uncertainity} = \frac{\text{absolute uncertainity}}{\text{measurement}} \times 100$$

$$= \frac{0{\cdot}2\ ^\circ c}{37{\cdot}5\ ^\circ c} \times 100 = 0{\cdot}53\%$$

Thus, the result is reported as $37.5 \pm 0.53\%$

22. The percentage in uncertainty of a measurement is 0.1%. The measured value is 8.314. Calculate the absolute uncertainty.

Solution:

$$\%\,\text{uncertainity} = \frac{\text{absolute uncertainity}}{\text{measurement}} \times 100$$

$$\text{Absolute uncertainty} = \frac{\%\,\text{uncertainty} \times \text{measurement}}{100}$$

$$= \frac{0{\cdot}1 \times 8{\cdot}314}{100} = 0{\cdot}008$$

2.3 Precision and Accuracy in Measurement

Precision tells us how closely multiple measurements of the same thing are to one another.

Accuracy is the closeness of the measurement to the **true or accepted value**

Example:

The true value of the density of water at 4°c and one atmosphere is 1g/mL. Two students X and Y determined the density of water at 4°C and 1atm and obtained the following values in g/mL.

Student	Trial						Average
	1	2	3	4	5	6	
X	0.91	0.92	0.90	0.88	0.97	0.90	0.90
Y	1.10	0.99	1.01	0.97	0.91	1.03	1.00

a) Which set of data is more precise?

b) Which set of data is more accurate?

Solution:

a) Student X has made more precise measurement than student Y.

b) Student Y has made more accurate measurement than student X.

Error = | True value − measured value |

There are two types of errors linked to precision and accuracy. These are:

Random errors - vary in direction and magnitude

- make the measurement less precise
- arise mostly from limitations in the instrument.

Systematic errors - lie in one direction of the true value, that is all above or below the true value.

- affect the result in a particular direction
- affect the accuracy of measurement
- arise from defects in the instrument and personal errors.

Illustrative Examples

23. The difference between an accepted value and an experimental value is called a (an) _____.

A) error C) percent error

B) measured value D) precision

Solution:

- Accepted value means true value
- Experimental value means measured value
- The difference between these values under absolute value is called an error.

Answer: A

24. The degree of reproducibility among the set of measurements is:

A) Accuracy C) Error

B) Precision D) Uncertainty

Solution:

The precision of a set of measurements refers to the degree of reproducibility among the set.

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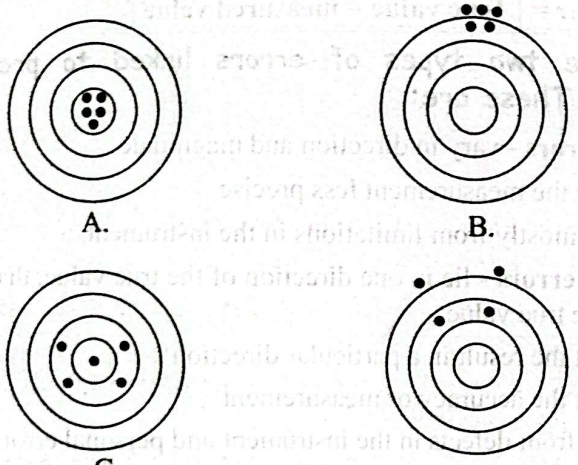
Answer: B

25. Four students used targets A, B, C and D while practicing with a rifle. The results are show below. How could you explain the results based on precision and accuracy?



Solution:

- Target A represents high accuracy and high precision
- Target B represents high precision and low accuracy.
- Target C represents low precision and high accuracy.
- Target D represents low precision and low accuracy.

26. The true mas of Aspirin table is 0.370g. Three students are asked to determine the mass of the tablet three times. The results in grams are given below.

Trial	Student 1	Student 2	Student 3
1	0.335	0.357	0.369
2	0.331	0.375	0.373
3	0.333	0.338	0.371

a) Determine the average value for each student

b) Which student has made precise but unaccurate measurement?

c) Which student has made both accurate and precise measurement?

d) Which student has made neither precise nor accurate measurement?

e) Identify the errors made by the students

Solution:

a) $\text{Student1} = \frac{0\cdot335+0\cdot331+0\cdot333}{3} = 0\cdot333$

$\text{Student2} = \frac{0\cdot357+0\cdot375+0\cdot338}{3} = 0\cdot357$

$\text{Student3} = \frac{0\cdot369+0\cdot373+0\cdot371}{3} = 0\cdot371$

b) Student 1 because the individual trials are close to each but the average is far away from the true value.

c) Student 3 because the trials are close to each other and the average is close to the true value

d) Student 1 has made systematic error while students 2 and 3 have made random error.

27. Four students measured the mass of gold whose true mass is 18.0g. Their results are 18.19g, 17.4g, 19.0g and 17.8g.

a) What is the best estimate for the mass of a piece gold.

b) Explain whether the results are precise or accurate.

Solution:

a) The best estimate value is the average value of the four students.

$\text{Average value} = \frac{18\cdot19+17\cdot4+19\cdot0+17\cdot8}{4} = 18\cdot075 \approx 18\cdot10$

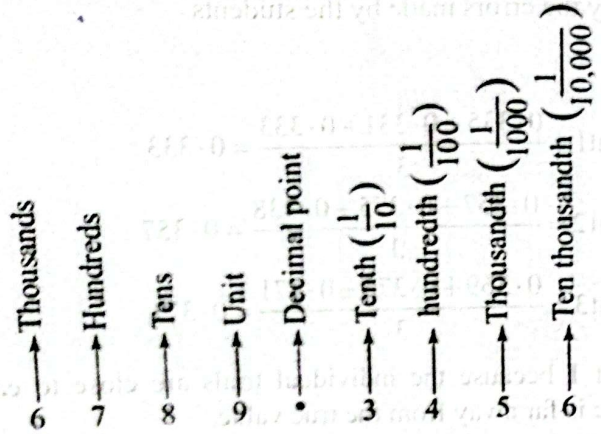
b) The average value is close to the true value. Thus, the measurement is accurate.

Decimal Place and Rounding off

A decimal place is the number of digital to the right of the decimal point.

A decimal point (•) is the point that separates the fraction part from the whole number part.

Consider the example below



Rounding off is finding a number that is closest to a given number but with fewer digits

Rules for Rounding of a number

1. If the digits to be removed is from **0 to 4**, the preceding digit remains same

   Example:

   Round off 7.89345 to two decimal place

   Answer: 7.89

2. If the digit to be removed is from 6 to 9 add one to preceding digit.

   Example:

   Round off 8.798679 to three decimal place.

   Answer: 8.799

3. If the digit to be removed is 5, the preceding digit is unchanged if it is even and increases by one if it is odd.

   Example:

   a) Round off 8.735 to two decimal place

   Answer: 8.74

   b) Round off 8.725 to two decimal place.

   Answer: 8.72

Note: If 5 is followed by zeros only, the left most digit is

unchanged.

- If 5 is followed by non-zero digits, the left most digit is increased by one

Example:

21.6500 → 21.6

21.6543 → 21.7

Numbers

Exact numbers	**Inexact numbers**
➥ Result from counting or defintions	➥ Obtained from measurement
➥ are abosuletly accurate	➥ are not accurate beacuse of uncertainty
Example	**Example**
→ The number of days in a week	→ The length of a table
→ The number of pages in this book	→ The mass of a coin
	→ The volume of a cup of tea

Significant figures are the meaningful digits in a reported number.

- The left digit in a measured number is referred to **uncertain digits** and the uncertainty associated with a measured number is generally considered as **± 1** in the place of the last digit.

For example: When we place the width of a card to be 2.5cm, we are implying that its width is **2.5 ± 0.1cm**

If the width of the card is 2.45cm, its width is **2.45 ± 0.01cm**

Rules to Determine Number of Significant Figures (s.f)

1. All non - zero digits are significant

   Example: 2395, s.f = 4

2. Zeros located between non - zero digits are significant

   Example: 507, s.f = 3

3. Zeros to the left the non - zeros digits are not significant.

   Example: 0.00035, s.f = 2

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4. Zeros to the right of the last non - zero digit are significant if the number contains a decimal point.

Example: 5.6000 has 5 s.f

➥ Zeros to the right of the last non-zero digit in a number without decimal point may or may not be significant. To avoid ambiguity, use scientific notation

Scientific notation is a way of expressing large or small numbers as a factor of the power of 10.

➥ It can be expressed $y \times 10^x$, where y is 1 to 10 excluding 10 and x is an integer

➥ When we shift the decimal point to the right, the power is negative and to the left the power becomes positive.

Example:

$0.00006234 = 6.34 \times 10^{-5}$

$564000000 = 5.64 \times 10^{8}$

Consider the number 500 without scientific notation, the number of significant figures is ambiguous.

We may write it with different significant figures without changing its magnitude.

$500 = 5 \times 10^2$ 1 s.f

$= 5.0 \times 10^2$ 2 s.f

$= 5.00 \times 10^2$ 3 s.f

$= 5.000 \times 10^2$ 4 s.f

: :

Calculations with Significant Figures

1. **In addition and subtraction:** the answer cannot have more digits to the right of the decimal point than any of the original numbers.

Example:

103.50 ⟶ two digits after the decimal place

$+0.316$ ⟶ three digits after the decimal place

102.816 ⟶ round to 102.82

2. **In multiplication and division,** the number of significant figures in the product or quotient is determined by the original number that has the smallest number of significant figures

Example:

$1.4 \times 6.782 = 9.4948 \rightarrow$ round to 9.5

$68.457 \div 5.35 = 12.795700935 \rightarrow$ round to 12.8

Illustrative Examples

28. How many significant figures are there in 0.006000?

A) 1 B) 6 C) 4 D) 5

Solution:

Zeros before 6 are not significant but zeros after 6 are significant due to the presence of decimal point.

Answer: C

29. All of the following are expressed with inexact numbers. **EXCEPT**

A) The thickness of a wire C) The area of a coin

B) The length of a ruler D) The number of days in a week.

Solution:

The number of days in a week are known exactly and expressed with exact number.

Answer: D

30. Determine the number of significant figures

a) 5.0680 c) 875 e) 0.0005

b) 10.00 d) 800 f) 2.008×10^6

Solution:

a) 5 c) 3 e) 1

b) 4 d) ambiguous f) 4

31. Solve the following and write your answer to the correct significant figures.

a) 317.5mL + 0.675mL c) 8.678mm + 10.251mm – 5.0mm

b) 47.80L – 2.075L

Solution:

a) 318.2ml b) 45.72ml c) 13.9mm

32. Determine the answer with the correct significant figure.

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a) 13.5g ÷ 45.18L

b) 6.25cm × 1.1750cm

c) $\frac{2\cdot673m \times 0\cdot82m}{126m}$

Solution:

a) 0.299g/L b) 7.34cm² c) 0.017m

33. Write the following numbers in scientific notation based on the given significant figures

a) 687951347 (2 s.f) c) 7984.6354 (1 s.f)

b) 0.00 000 000 6758 (3 s.f)

Solution:

a) 6.9×10^8

b) 6.76×10^{-10}

c) The fraction part is very small compared to the whole number part. As a result, you should focus on the whole number part.

- 8×10^3

34. What is the purpose of scientific notation?

Solution:

- To avoid ambiguity in determining significant figures that ends with zeros.
- To write a very large or small number easily.

2.4 The Scientific Method

Scientific method is logical, careful and systematic approach to solve scientific problems.

Steps in Scientific Method

1. Observation and Formulation of Question

- It is looking through carefully the nature of an event
- There are two types of observations. These are

i. **Quantitative Observation** - is numerical observation (has numbers with it).

Example: A book weighs 100 grams

Water boil at 100°C.

ii. **Qualitative observation** - is non-numerical (has no numbers with it)

Example: The sky is blue.

Honey is sweet.

2. Hypothesis and Data Collection

- **Hypothesis** is tentative explanation for an observation.
- Hypothesis is useful if it can be used to make predictions

3. Experiment (Testing hypothesis)

- It is examining facts about hypothesis
- It is required to check whether the hypothesis agrees or disagrees the observation made in the real world.

4. Analysis and conclusion

- It involves a statement or mathematical equation that summarizes facts obtained from many experiments.
- It serves as a convenient means of storage for large quantities experimental facts.

Illustrative Example

35. What is the first step in scientific method?

A) Making observations

B) Formulating hypothesis

C) Performing an experiment

D) Predicting the result of can experiment

Solution:

The first step in scientific method is observation and formulation of question.

Answer: A

36. Inorder to advance to the next step, a hypothesis should:

A) be obviously accepted by many people

B) be repeatedly confirmed by experimentation

C) be fully functional experiment

D) report the past experience

Answer: B

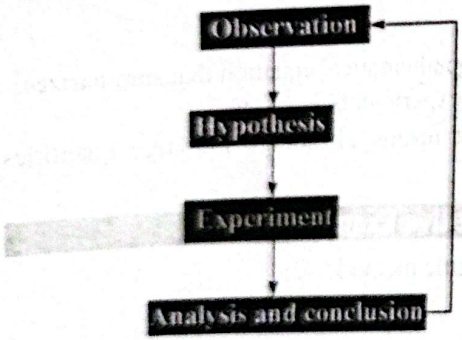
37. Classify the following as qualitative or quantitative observations
   a) The leaves of most trees are green in summary
   b) The people consists of over 95% water
   c) Chemistry is an easy subject
   d) I got 90% is my last chemistry examination
   e) My hair is black

Solution:

- a, c and e are qualitative observation.
- b and d are quantitative observation.

38. Write the steps in scientific method using diagram.

Solution:



2.5 Some Experimental Skills in Chemistry

Chemistry is an experimental science that relates theoretical study with practical activities.

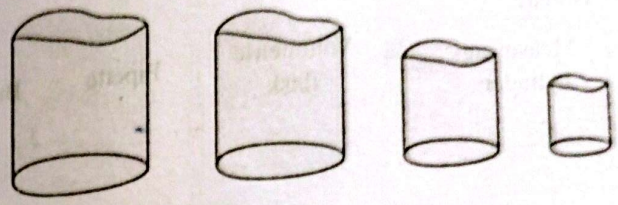
Chemistry Laboratory Apparatus

- Laboratory apparatus are classified into three based on their purposes These are

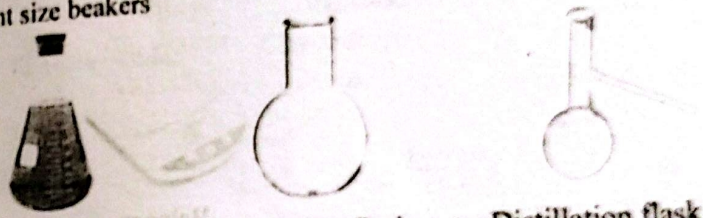
1. **Reaction vessels** are used for the occurrence of chemical reactions.

   Example: Beakers

- Flasks
- Test tubes
- Evaporating dish
- Crucible

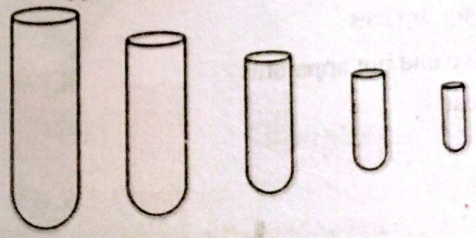


a) Different size beakers



Conical flask Florence flask Distillation flask

b) Different Types flasks



c) Different size test tubes

Evaporating dish Crucible and Cover Watch glass

d) Other reaction vessels.

2. **Measuring equipments** - are used to measure amount of substances.

Example:

- Balances
- Measuring cylinder
- Pipette
- Burette
- Volumetric flask
- Thermometer
- pH meter

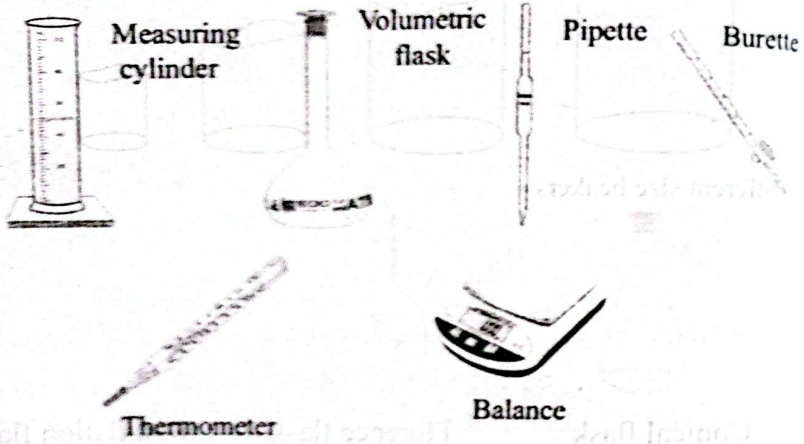


Fig 2.6 Measuring Apparatus

3. Supporting and heating devices

- used to hold sensitive and hot apparatus
- used as source of heat

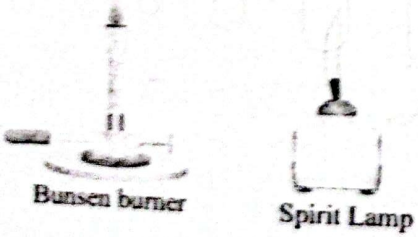


Fig 2.7 Heating Devices

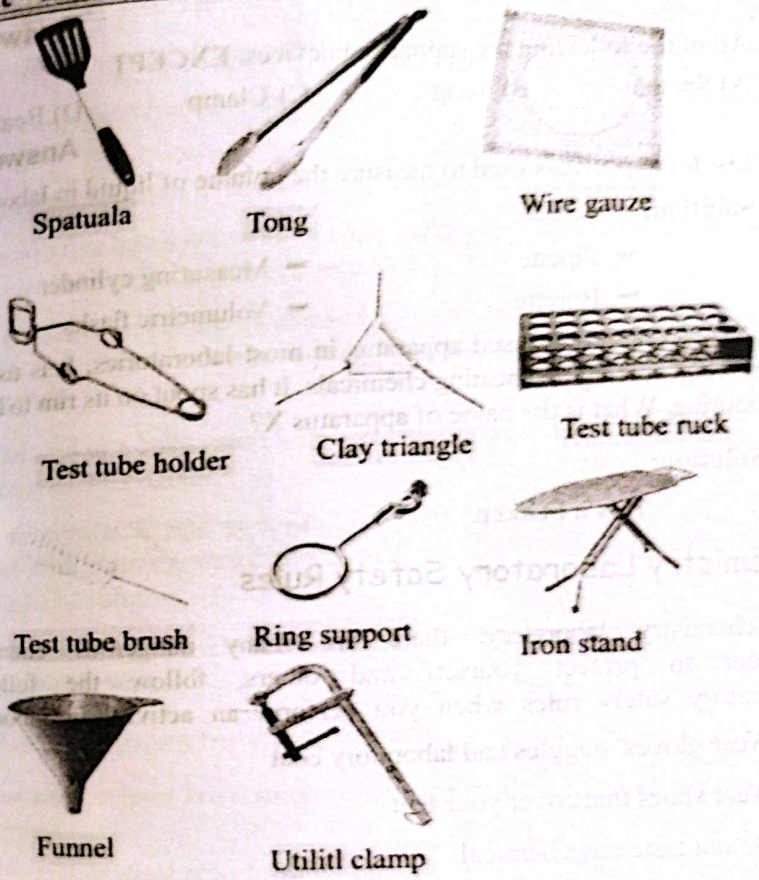


Fig 2.8 Commonly used Supporting Devices

Illustrative Examples

39. Which of the following is used as reaction vessel?

A) Test tube
B) Iron stand
C) Thermometer
D) Bunsen burner

Solution:

Iron stand - supporting device
Thermometer - measuring apparatus
Bunsen burner - used as source of heat

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Answer: A

40. All of the following are supporting devices. **EXCEPT**

A) Spatula B) Tong C) Clamp D) Beaker

Answer: D

41. List four apparatus used to measure the volume of liquid in laboratory

Solution:

- Pipette
- Measuring cylinder
- Burette
- Volumetric flasks

42. X is a commonly used apparatus in most laboratories. It is used for mixing, stirring and heating chemicals. It has spout on its rim to help in pouring. What is the name of apparatus X?

Solution:

It is a beaker.

Chemistry Laboratory Safety Rules

In chemistry laboratory, there are many dangerous chemicals Inorder to protect yoursetf and others, follow the following laboratory safely rules when you perform an activity in laboratory.

1. Wear gloves, goggles and laboratory coat
2. Wear shoes that cover your feet.
3. Do not taste any chemical.
4. Tie back lose hair.
5. Do not suck the solutions, in the pipette by mouth, rather use a rubber suction bulb or other device to fill a pipe.
6. Do not eat or drink anything in laboratory.
7. To dilute concentrated acids, add acid to water **(AW!)** not water to acid **(WA!)**
8. Never look directly into a test tube
9. Point test tubes that are being heated away from you and others.
10. Carry out only allowed experiments

11. Read the lables of the reagent bottles carefully to make sure that you are using the right chemicals
12. When lighting Bunsen burner, first light the match stick then turn on the gas
13. Know the location of fire extinguisher and instruments.
14. If corrosive bases are swallowed, weak acids such as **acetic acid or lemon acid juice** is administered.
15. If corrosive acids are swallowed by accident, administer weak bases such as magnesium hydroxide **($Mg(OH)_2$), aluminium hydroxide ($Al(OH)_3$).**
16. If an acid or a base enters your eye, wash with water repeatedly and then consult a doctor.
17. If concentrated acid is spilled on your body, first wash the affected part with running water and then wash with 10% sodium carbonate (Na_2CO_3) solution.
18. Before leaving the laboratory, clean your apparatus, work surface and your hands well.

Writing a Laboratory Report

A laboratory report is a written composition of the results of an experiment

To write a laboratory report follow the steps below

1. Title _____
2. Objective _____
3. Materials and chemicals _____
4. Procedure _____

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5. Observation _____

6. Results and discussion _____

7. Conclusion _____

Laboratory Report Answer the Following questions.

- What did you do?
- Why did you do it?
- How did you do it?
- What happened?

Illustrative Examples

43. Which of the following is not true about safety rules?

A) Wear approved eye protection
B) Do not smoke in the laboratory
C) Add water to acid to dilute an acid.
D) Report all injuries to teacher or laboratory technician.

Answer: C

44. State some laboratory rules that are not stated in this book.

Solution:

- Never remove chemicals from the laboratory
- Never return extra chemicals into the reagent bottles.
- Add boiling chips to liquids to be boiled.

- Never place chemicals directly on the pan balance, use filter paper.

45. Write a short laboratory report of the experiment of burning magnesium ribbon to produce basic oxide.

Solution:

Title: Burning of magnesium ribbon

Objective: To prepare basic oxide from burning of magnesium ribbon.

Materials and chemicals:

Magnesium ribbon, Bunsen burner, test tube or crucible, tong, water, litmus paper

Procedure: Turn on the Bunsen burner, hold the magnesium ribbon with tong and burning it.

Observation: A bright light was observed when magnesium burnt in air. The white ash formed was collected in crucible.

Result and discussion: At the end of burning, white ash was formed and dissolved in water. When the solution was checked with indicators it turned the litmus paper to blue.

Conclusion: The white ash formed was known to be basic oxide since it changed the red litmus paper to blue.

Solved Problems

46. List the seven basic SI units.

Solution:

- Meter (m)
- second (s)
- Candela (Cd)
- kilogram (kg)
- mole (mol)
- Kelvin (K)
- Ampere (A)

47. What are the basic SI units involved to device the SI unit of force?

Solution:

Force = mass × acceleration

$F = m \times a$

$1N = 1kg \cdot m/s^2$

Thus, the basic SI units used to derive Newton (N) are kilogram (kg), meter(m) and second (s).

48. Relate the different units of length.

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Solution:

1km = 10^3m = 10^5cm = 10^6mm = 0.62 mile

1m = 10dm = 100cm = 1000mm

1m = 1.094yard

1inch = 2.54cm = 0.0254m = 25.4mm

1foot = 12 inch

1 yard = 3 feet

1 mile = 5280 feet

49. Relate the units of mass, kg, mg, pound and meteric tone

Solution:

- 1kg = 10^3g = 10^6mg
- 1 ounce = 0.02835kg
- 1 pound = 0.4536kg
- 1 metric tone = 1000kg

50. Explain the relationship between units of volume (L,mℓ, cm^3 and dm^3)

Solution:

$1m^3 = 10^3\ell = 10^6cm^3 = 10^6m\ell = 10^9mm^3$

$1\ell = 1dm^3$

$1m\ell = 1cm^3$

$1\ell = 1.06$ quarts

51. Bromine is a reddish- brown liquid. Calculate its density if 586 g of it occupies 188ml of volume.

Solution:

$$\rho = \frac{m}{V} = \frac{586g}{188mL} = 3{\cdot}12g/mL$$

52. Oil floats on the surface of water but mercury sinks. Explain why?

Solution:

Oil is less dense than water but mercury is denser than water.

53. Calculate the density of a rectangular block of metal whose length is 8.5cm, width is 1.85cm, height is 3.24cm and mass 65.672g.

Solution:

$$\rho = \frac{m}{V} = \frac{m}{l \times w \times h} = \frac{65{\cdot}672g}{8{\cdot}5cm \times 1{\cdot}85cm \times 3{\cdot}24cm} = 1{\cdot}3g/cm^3$$

54. Aluminium has a density of 2.70g/ml. What mass of a piece of aluminium occupies a volume of 0.01L?

Solution:

First convert the unit of volume into ml.

1L = 1000ml

0×01L = x

$$x = \frac{0 \times 01L \times 1000ml}{1L} = 10ml$$

$$\rho = \frac{m}{V}$$

$$m = \rho \times V = 2{\cdot}70\frac{g}{ml} \times 10ml = 27g$$

55. Oxygen freezes to a solid at −218°C. Convert this temperature into

a) °F b) K

Solution:

a) $°F = \frac{9}{5} \times {}^0C + 32$

$= \frac{9}{5} \times (-218) + 32 = -360{\cdot}4\,°F$

b) $K = {}^0C + 273$

$= -218 + 273 = 55K$

56. How many second are there in a solar year (365.24 days)

Solution:

1day = 24hours = 1440 minutes = 86,400 seconds

$$\therefore\ 365{\cdot}24days = 86,400\frac{s}{day} \times 365{\cdot}24days = 3{\cdot}156 \times 10^7\ sec.$$

57. Convert the following

a) 0.072 m^3 to dm^3 b) 6A° to μm c) 6.72×10^{-7}m to nm

Solution:

a) $1m^3 = 1000dm^3$

$0{\cdot}0712m^3 = x$

$$x = \frac{0{\cdot}0712\,m^3 \times 100dm^3}{1\,m^3} = 71{\cdot}2dm^3$$

Thus, $0{\cdot}0712m^3 = 71{\cdot}2dm^3$

b) $$\frac{1A^0}{1\mu m} = \frac{10^{-10}\,m}{10^{-6}\,m} = 10^{-4}$$

$\Rightarrow 1A^0 = 10^{-4}\mu m$

$\therefore\ 6A^0 = 6 \times 10^{-4}\mu m$

c) $6{\cdot}72 \times 10^{-7}m = 6{\cdot}72 \times 10^{-7}(10^9 nm) = 672nm$

Reason : $1m = 10^9 nm$

58. A thermometer reading of 37.5°C has an absolute uncertainty of 0.1°C. Calculate the percent uncertainty and report the result of measurement interms of absolute uncertainty and percent uncertainty.

Solution:

$$\%\,\text{uncertainty} = \frac{\text{absolute uncertainty}}{\text{measurement}} \times 100$$

$$= \frac{0{\cdot}1\,°C}{37{\cdot}5\,°C} \times 100 = 0{\cdot}27\%$$

Result : $37{\cdot}5°C \pm 0{\cdot}1°C$

$37{\cdot}5°C \pm 0{\cdot}27\%$

59. What is the relationship between precision and uncertainty of a measurement?

Solution:

The lower the uncertainty, the more precise is the measurement.

60. Students A and B determined the mass of an object having an actual mass of 10.24g using two different balances. Each student has made three trials as tabulated below.

Student	Trail 1	Trial 2	Trial 3	Average
A	10.26g	10.24g	10.20g	10.23g
B	10.20g	10.18g	10.22g	10.20g

a) Indicate the type of errors, systematic or random error involved in the measurements of students A and B.

b) Show which of the measurements is more precise.

c) Which student has made accurate measurement?

Solution:

a) Student A has made random error because one value is below and the other one is above the true value.

Student B's measurement involves systematic error because all the trials are below the true value,

b) Student B's measurement is more precise because individual trial is close to each other.

c) Student A has made more accurate measurement because the average value is closer to the true value.

61. Determine the number of significant figures in each of the following measurements.

a) 0.0080L
b) 60,156 tons
c) 6 kg
d) 9.50×10^4cm
e) $10.00m^3$
f) 1000m

Solution:

a) 2
b) 5
c) 1
d) 3
e) 4
f) ambiguous

62. Carry out the following operations in the correct number of significant figures

a) 7.310 km ÷ 5.70 km

b) $(3.26 \times 10^{-3}mg) - (7.88 \times 10^{-5}\ mg)$

c) $(4.02 \times 10^4 m) + (7.74 \times 10^7\ dm)$

d) $(1.0 \times 10^4 m) \times (9.9 \times 10^6 m)$

e) $(7.95 \times 10^4 g) \div (2.5 \times 10^4\ m\ell)$

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Solution:

a) 1.28

b) Before you subtract make powers of 10 equal

$= (3.26 \times 10^{-3}mg) - (0.0788 \times 10^{-3}mg)$

$= 10^{-3}mg\,(3.26 - 0.0788) = 3.18 \times 10^{-3}mg$

c) First convert into similar unit.

$= (4.02 \times 10^{5}dm) + (7.74 \times 10^{7}dm)$

$= (4.02 \times 10^{5}dm) + (774 \times 10^{5}dm)$

$= 10^{5}dm\,(4.02 + 744)$

$= 778.02 \times 10^{5}dm$

$= 7.78 \times 10^{7}dm$ or $7.78 \times 10^{6}m$

d) $9.9 \times 10^{10}m^2$

e) 3.2g/ml

63. A measuring cylinder has a precision of ± 0.01ml. A liquid sample that has a volume of 10ml is measured by measuring cylinder. How many significant figures should be reported for the measurement?

Solution:

A precision ± 0.01ml implies there are two decimal places after the whole number 10.00mL. As a result, the number of significant figures is 4.

64. Write the following numbers in scientific notation with the given number of significant figures.

a) 9,998,000,000,000 (3 s.f)

b) 0.000,000,000,000,000,00895 (2 s.f)

Solution:

a) 1.00×10^{13} b) 9.0×10^{-18}

65. Convert $1A^{0^3}$ into cm^3 and nm^3.

Solution:

$$1A^{0^3} = \left(10^{-10}m\right)^3 = \left(10^{-10} \times 100cm\right)^3$$

$$= 1 \times 10^{-24}cm^3$$

$$1A^{0^3} = \left(10^{-10}m\right)^3 = \left(10^{-10} \times 10^{9}nm\right)^3 = 10^{-3}nm^3$$

Supplementary Exercise

I. True/False Items

1. Chemist use measurement to compare the properties of different substances.
2. Density is a derived quantity obtained from basic quantities mass and volume.
3. Heat flows non- spontaneously from cold body to hot body.
4. Systematic uncertainty causes the value to be too large or too small.
5. The number of pages in this book can be expressed with inexact number.
6. In addition and subtraction of significant figures answer can have more digits to the right of the decimal point of original numbers.
7. The aim of all scientific methods is different and involve different steps.
8. Chemistry is an experimental science that requires careful and accurate measurement.
9. Intensive properties are physical properties that depend on amount of substance.
10. The product of 3567 and 2.5 has two significant figures.

II. Matching Items

11. Ato	A) 10^{-6}
12. Femto	B) 10^{-9}
13. Hecto	C) 10^{12}
14. Tera	D) 10^{-18}
15. Pico	E) 10^{-15}
16. Exa	F) 10^{-12}
17. Deca	G) 10^{15}
18. nano	H) 10^{18}
	I) 10

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19. Mega — J) 10^{-3}
20. Giga — K) 10^{6}
L) 10^{9}
M) 10^{2}
N) 10^{3}

III. Multiple Choice Items.

21. Which of the following is used to measure length?
A) balance C) burette
B) thermometer D) meter stick

22. Which of the following is measured by an indirect method?
A) Volume of a cup water
B) The length of a table
C) The temperature of the classroom
D) The mass of an atom of hydrogen

23. Which of the following has no unit?
A) Amount of substance C) Luminous intensity
B) Specific gravity D) Weight

24. Which of the following is an extensive property?
A) colour C) area
B) conductivity D) boiling point

25. Which of the following is not allowed in the chemistry laboratory?
A) Wearing any type of cloth and shoe
B) Knowing the hazards of the materials being used.
C) Reading the lables of the reagent bottles carefully.
D) Using glycerin when inserting glass tubing into rubber stoppers.

26. Which of the following has three significant figures?
A) 530 B) 0.7130 C) 0.003 D) 4.01

27. A blood sample analysis showed a total of 0.00467g lead. What is the result if it is expressed in micrograms?
A) $4.67 \times 10^{3}\mu g$ C) $4.67 \times 10^{6}\mu g$
B) $0.467\mu g$ D) $4.67 \times 10^{4}\mu g$

28. Which of the following is true?
A) A substance with density 2.7g/ml floats on water
B) A scientific notation for 0.00021 is 2.1×10^{-4}.
C) One meter has about the same length as a yard.
D) The prefix kilo means multiplied by 10^{-3}.

29. Which of the following has the lowest uncertainty?
A) 6.0 B) 6.00 C) 6.000 D) 6.0000

30. All of the following are supporting devices in chemistry laboratory. EXCEPT
A) test tub B) tong C) iron stand D) tripod

31. The number of significant figures in the result of 6.7245 divided by 10.3 is _____
A) 5 B) 3 C) 4 D) 2

32. Which of the following express the one billionth of in four significant figures?
A) 1×10^{-9} B) 1.0×10^{-9} C) 1.00×10^{-9} D) 1.000×10^{-9}

33. A measurement gave a mass of 0.65kg. This is the same as:
A) 6.5g B) 65g C) 650g D) 4500g

34. What is the mass of 30.0ml of a liquid that has a density of 1.60g/ml.
A) 18.8g B) 48.0g C) 31.6g D) 53.3g

35. A conversion factor, like the definition 1km = 1000m is
A) exact C) cannot be predicted
B) inexact D) subject to individual opinion

36. Which one of the following is wrong?
A) $1L = 1dm^3$ C) $1kg|L = 1g|cm^3$
B) $1mL = 10cm^3$ D) $1g|L = 1kg|m^3$

37. Which of the following is the SI unit of area
A) m^3 B) mm^2 C) m^2 D) cm^2

38. Which one is the largest unit of temperature?

A) °C B) K C) °F D) None

39. _____ is tentative explanation for an observation.

A) Hypothesis C) Theory

B) Experiment D) Scientific law

40. 86dL = _____ L

A) 86 B) 860 C) 8.6 D) 0.86

I. Short Answer Items

41. Convert the following

a) 770mg = _____ g

b) 350μs = _____ s

c) 550nm = _____ mm

42. Convert the following temperatures as indicated.

a) −270°C to K b) −300°F to K c) 500K to °F

43. How many seconds are there in 10 minutes?

44. Carry out the following conversions.

a) 500mL to dm^3 c) $125A^{0^3}$ to nm^3

b) 850 L to m^3

45. Perform the indicated operations and give answers with the proper number of significant figures.

a) 48.2 m + 3.82m + 48.439m d) $73.0 \times 1.340 \times 0.46$

b) 551g − 15.46g e) $\frac{22.61 \times 0.0587}{135 \times 28}$

c) 15.44mL − 9.1ml + 105mL

46. Calculate the number of mole of 171g of sucrose ($C_{12}H_{22}O_{11}$)

47. State the number of significant figures in the following quantities.

a) 6.00g of gold

b) 0.000456 sec

c) 6.0001×10^6 mL of water

48. Round of the following numbers to the given significant figures.

a) 8.345 (3 s.f) b) 8.435 (3 s.f) c) 2.9856 (2 s.f)

49. The _____ of a measurement describes how well repeated measurements of a value agree with one another.

50. _____ is determined by the agreement between the measured quantity and the correct value.

Answer key to Supplementary Exercise

Part I

1. True	3. True	5. False	7. False	9. False
2. False	4. True	6. False	8. True	10. True

Part II

11.D 12.E 13.M 14.C 15.F 16.H 17.I 18.B 19.K 20.L

Part III

21.D	24.C	27.A	30.A	33.C	36.B	39.A
22.D	25.A	28.B	31.B	34.B	37.C	40.C
23.B	26.D	29.D	32.D	35.A	38.B	

Part IV

41.a. 0.770g

b. 0.000350sec = 3.50×10^{-4}sec

c. 0.000550 mm = 5.50×10^{-4} mm

42.a. 3K b. 88.6 K c. 440.6°F

43. 600 seconds

44.a. 0.500dm^3 b. 0.850m^3 c. 0.125nm^3

45.a. 100.5m b. 536g c. 536g d. 45 e. 0.00035 or 3.5×10^{-4}

46. 0.5mol

47.a. 3 b. 3 c. 5

48.a. 8.34 b. 8.44 c. 3.0

49. Precision

50. Accuracy

UNIT 3

Structure of the Atom

3.1 Historical Development of the Structure of the Atom

- ☑ Early thoughts concerning atoms were proposed by the Greek philosopher, **Democritus** about 400B.C
- ☑ He suggested that the world was made of two things **empty space and tiny particles he called " atoms"**
- ☑ The word atom comes from the Greek word **atomos,** meaning indivisible.
- ☑ The theory of Democritus was very general and was not supported by experimental evidences.
- ☑ **Aristotle** was the opponent of Democritus that did not believe in the existence of atom.
- ☑ He believed that matter was continuous and his teachings were accepted unit the 17th century.

Table 3.1 Comparison Between Continuous and Discontinuous Theory

Discontinuous or Particle Theory	Continuous Theory
➥ By Democritus	➥ By Aristotle
➥ Division has limit	➥ Division has no limit
➥ Matter was not continuous	➥ Matter was continuous
➥ Believed in the existence of atom	➥ Did not believe in the existence of atom

Illustrative Examples

1. The first atomic theory was proposed by the Greek philosopher _____.

A. Arstotle C. Democritus

B. Plato D. Leucippus

Solution

Plato - introduced the atomic theory in which ideal geometric forms serve as atoms.

Aristotle - Believed that there were only Five Elements:

- ➥ **Air** - which was light
- ➥ **Earth** - which was cool and heavy
- ➥ **Water** - which was cold
- ➥ **Fire** - which was hot
- ➥ **Aether** - which he viewed as divine substance which made up of the stars and planets.

Leucippus

- ➥ Little is known about him
- ➥ He was the teacher of Democritus.

Answer: C

2. Who was know to be the teacher of Aristotle and student of Socrates?

A. Plato B. Democritus C. Leucippus D. Dalton

Answer: A

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3. What are the four elements used to make materials based on the thoughts of Greek philosopher, Empedocles ?

Solution

Air, earth, water and fire

4. The Democritus theory of atoms generally consists of 5 basic principles. What are they?

Solution

- Every matter is composed of atoms.
- Atoms are in solid form but we cannot see them
- Atoms cannot be destroyed.
- Atoms have different properties in terms of size, form, weight, position and type of arrangement.

5. What was the shortcoming or draw back of the early Greek philosophers?

Solution

Their reasonable and logical understanding of the atomic theory was not supported by experiment and observations.

3.2 Fundamental Laws of Chemical Reactions

Scientific law is a concise verbal or mathematical statement that summarizes facts that come from many experiments.

There are three fundamental laws in chemical reaction. These are

1. Law of conservation of mass
2. Law of definite proportion
3. Law of multiple proportion

1. Law of Conservation of Mass

- It is also called law of indestructibility of matter.
- It was expressed by **A. Lavoiser** in 1774 .
- He proposed that inordinary chemical reactions, matter can be changed in many ways, but it cannot be created or destroyed

- According to law of conservation mass, in balanced chemical reaction;

Mass of reactants = mass of products.

Example:

$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

$28g + 3 \times 2g = 2 \times 17g$

$34g = 34g$

2. Law of Definite Proportion (law of constant composition)

It was proposed by the French chemist **Joseph Proust.**

He stated that the elements in a compounds have definite proportion by mass.

Example:

Water is made of hydrogen and oxygen. The ratio of the mass of hydrogen to oxygen in any sample of pure water is always the same. No matter where the sample is obtained, how it is obtained or how large it is, the ratio of the mass of hydrogen to oxygen never changes.

- H_2O, the mass ratio of H to O is 1: 8

3. Law of Multiple Proportion (Daton, 1803)

The statement of this law is the ratio of the masses of one element that combine with a constant mass of another element can be expressed in small whole numbers. Consider table 3.2 below to understand it.

Table 3.2 The Mass Ratio of Oxygen in CO and CO_2

	Proportion of carbon to oxygen		
	Mass of C in 1 mole	Mass of O in 1 mole	Ratio of O masses combined with constant mass of carbon
Carbon mono oxide, CO	12g	16g	$\frac{16}{12} \div \frac{16}{12} = 1$

Carbon dioxide, CO_2	12g	32g	$\frac{32}{12} \div \frac{16}{12} = 2$

There fore, the mass, ration of oxygen with fixed mass of carbon is 1:2

Example

SnO and SnO_2, H_2O_2 and H_2O, N_2O, NO, N_2O_3, NO_2 and N_2O_5 illustrate the law of multiple proportion.

Illustrative Examples

6. When 16kg of methane is burned with 64kg of oxygen, 44kg of carbon dioxide and 36g of water were produced. This an example of the law of

A. Equivalent proportions
B. Multiple proportions
C. Conservation of mass
D. Definite proportions

Solution

Mass of CH_4 + mass of O_2 = CO_2 of + H_2O of

16 g + 64 g = 44 g + 36 g

80 g = 80 g. This is true based on the law of conservation of mass

Answer: C

7. Nitrogen forms different oxides with oxygen upon the reaction conditions. This phenomenon can be summarized by the law of:

A. Conservation of mass
B. Combining Volumes
C. Multiple proportion
D. Constant composition

Solution

Two or more compounds formed from the same elements illustrate the law of multiple proportion.

Answer: C

8. "A compound always contains the same elements combined together in the same proportion by mass" This is according to the law of:

A. Multiple proportion
B. Conservation of mass
C. Conservative of energy
D. Definite proportions

Answer: D

9. Explain the law of multiple proportions by using the five oxides of nitrogen.

Solution

Mass of N = 14, mass of O = 16

Formula of oxide	Mass Ratio of N to O	Mass Ratio of O with Fixed mass of N
N_2O	$28 : 16 : = \frac{28}{28} : \frac{16}{28} = : 1 : \frac{4}{7}$	Dividing by the smallest $\frac{4}{7} \div \frac{4}{7} = 1$
NO	$14 : 16 : = \frac{14}{14} : \frac{16}{14} = : 1 : \frac{8}{7}$	$\frac{8}{7} \div \frac{4}{7} = 2$
N_2O_3	$28 : 48 : = \frac{28}{28} : \frac{48}{28} = : 1 : \frac{12}{7}$	$\frac{12}{7} \div \frac{4}{7} = 3$
NO_2	$14 : 32 : = \frac{14}{14} : \frac{32}{14} = : 1 : \frac{16}{7}$	$\frac{16}{7} \div \frac{4}{7} = 4$
N_2O_5	$28 : 80 : = \frac{28}{28} : \frac{80}{28} = : 1 : \frac{20}{7}$	$\frac{20}{7} \div \frac{4}{7} = 5$

Therefore, the ratio of oxygen in the five oxides of with fixed mass of nitrogen is 1 : 2 : 3 : 4 : 5

10. Who stated that the elements in a substances have definite proportion by mass?

Answer: Joseph Proust

11. Calculate the mole of 31.5g of nitric acid (HNO_3).
(Atomic Mass : H = 1, N = 14, O = 16)

Solution

$n = \frac{m}{M}$ m – given mass, M – molar mass

Molar mass = sum of total atomic mass of the elements in the compound in gram per mole (g/mol)

Total atomic mass = subscript × atomic mass

$\therefore M(HNO_3) = 1 \times 1 + 1 \times 14 + 3 \times 16 = 63$ g/mol

$$n = \frac{m}{M} = \frac{31.5\ g}{63\ g/mol} = 0.5\ mol$$

12. Describe the mass ratio of carbon to oxygen in carbon monoxide (CO)

(Atomic mass : C = 12, O : 16)

Solution

Mass ratio of C to O = $\frac{\text{total atomic Mass of C}}{\text{total atomic mass of O}}$

= 12 : 16

= Simplified as 3 : 4

13. What is the ratio of carbon to hydrogen to oxygen (by mass or weight) in 15g of glucose ($C_6H_{12}O_6$) ?

Solution

What ever the amount of glucose given, the ratio of C to H to O by weight will always be same. That is, the same as the simplified total atomic mass of the elements in the compound.

	C	:	H	:	O
$C_6H_{12}O_6$	6×12	:	$12 \times 12 \times 1$	:	6×16
	= 72	:	12	:	96

= Simplified as 6 : 1 : 8

3.3 Atomic Theory

Atomic theory is the study of the nature of atoms and the forces that hold them together.

Dalton's Atomic Theory

John Dalton, a British school master and chemist was known to be the founder of modern atomic theory. This is because his postulates were stated based on the three fundamental chemical laws.

The Main Postulates of Dalton's Atomic Theory are

1. All substances are made up of atoms
2. Atoms are indivisible and indestructible
3. All atoms of a given element are identical in mass and other properties
4. Atoms of different elements are different.
5. Atoms unite in definite ratio to form compounds.
6. A chemical reaction involves a change not in the atoms themselves but in the way atoms are combined to form compounds.

Note

- Postulate 2 and 3 were known to be incorrect by later experimental findings.
- Dalton proposed the symbol and formula of same elements and compounds respectively.

Table 3.3 Dalton's Symbols and Formula

Dalton's Symbol of Elements	Dalton's Formula of Compounds
Carbon	Alumina
Copper	Ammonia
Hydrogen	Carbonic Acid
Lead	Carbonic Oxide
Nitrogen	Olefiant
Oxygen	Potash
Phosphorus	Soda
Sulphur	Sulphuric acid
	Water

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Modern Atomic Theory

The modern atomic theory was developed because of the experimental works of many scientists. For example, the discovery of sub-atomic particles and isotopes.

It can be summarized as follows:

1. Elements are made up of small particles called atoms.
2. Atoms are divisible in to sub-atomic particles electron, proton and neutron
3. Atoms cannot be destroyed or cratered during ordinary chemical reactions.
4. Atoms of the same element have different masses.
5. Atoms of different element are different.
6. Atoms combine in small whole number ratio to form compounds.

Illustrative Example

14. Which of the following is not true about Daltno's atomic theory?

A. Atoms of different elements are different
B. Atoms are indivisible
C. Atoms of the same element have different masses.
D. Atoms combine in small whole number ratio to form compound.

Solution

According to Dalton, atoms of the same element have the same mass

Answer: C

15. All of the following are draw backs of Dalton's atomic theory. EXCEPT

A. Atoms are indivisible
B. Atoms of the same elements have the same mass
C. Atoms of different elements combine in simple whole number ratio to form compounds.
D. None

Answer: C

16. Write two modified postulates of Dalton.

Solution

➥ Atom is indivisible

➥ Atoms of the same element have same masses.

17. Write the Dalton's symbol for:

a. Hydrogen b. Oxygen c. Carbon

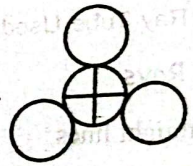
Solution

a. b. c.

18. Write the Dalton's formula for:

a. Water b. Sulphuric acid c. Olefiant

Solution

a. b.  c.

19. Write the advantages (merits) of Dalton's atomic theory.

Solution

➥ His theory explains the law of conservation of mass and law of multiple proportions.

➥ He recognized the difference between an element (atom) and a compound (molecule).

20. Based on the Dalton thought, what are the smallest particles of matter?

Answer: Atoms

3.4 Discoveries of the Fundamental Sub Atomic Particles and the Atomic Nucleus

1. Discovery of Electron

☑ J.J Thomson used a discharge tube to study the electron. A light potential difference was applied a cross the two electrodes **(anode and cathode)**

☑ When electricity was switched on a beam of rays emitted from the

cathode. These rays were called **cathodes rays** because they came from the cathode.

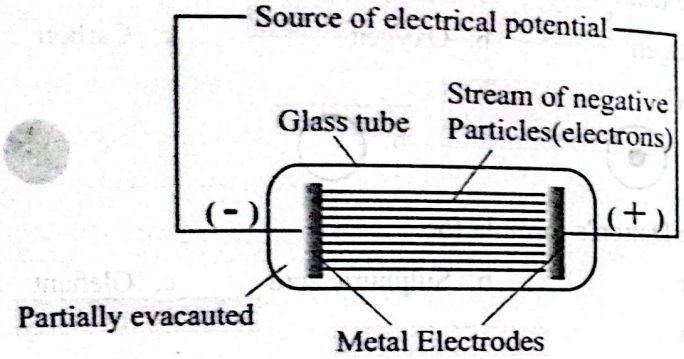


Fig 3.1 Cathode Ray Tube Used by Thomson

Properties of Cathode Rays

- They move in straight lines
- They have constant charge to mass ratio.
- They are deflected by both magnetic and electric fields.
- They ionize the gases
- They are deflected toward positive plate.
- They produce X-rays when allowed to fall on metals such as tungsten and copper
- They are found in all elements.
- They have penetrating power.

Note :-

(i) The charge of an electron was determined by **R. Millikan using oil - drop experiment.**

- The charge of an electron was found to be $-1.602 \times 10^{-19} C$

(ii) The charge to mass ratio (e^-/m) of an electron was determined by **J. J Thomson.** $e^-/m = 1.76 \times 10^8 C/g$

From charge and charge to mass ratio, it is possible to calculate mass of an electron (m_{e^-})

$$m_{e^-} = \frac{e^-}{e^-/m} = \frac{-1.602 \times 10^{-19} C}{-1.76 \times 10^{8} C/g}$$

$$m_{e^-} = 9.10 \times 10^{-28} g$$

$$m_{e^-} = 9.10 \times 10^{-31} kg$$

(iii) An electron has negative charge and the highest charge to mass ratio.

Illustrative Examples

21. Which of the following was determined by the Millikan's oil drop experiment?

A. Charge of an electron
B. Charge to mass ratio of an electron
C. Mass of an electron
D. Velocity of an electron

Solution

The **Millikan's oil - drop** experiment was used determine the charge of an electron.

Answer: A

22. Which one is not true a bout an electron?

A. It has the smallest mass
B. It has negative charge
C. It has the smallest charge to mass ratio
D. It is found in all atoms.

Answer: C

23. How could you calculate the mass of an electron?

Solution

From charge and charge to mass ratio of an electron.

24. Calculate the mass of Avogadro's number of electrons.

Solution

Avoagdro's number = 6.02×10^{23}

Mass of an electron = 9.1×10^{-28}g

Mass of $6.02 \times 10^{23} = x$

$$x = \frac{\text{Mass of } 6.02 \times 10^{23} \text{ electrons} \times 9.1 \times 10^{-28}\text{ g}}{\text{Mass of an electron}}$$

$$x = 4.48 \times 10^{-4}\text{ g}$$

➥ Mass of Avogadro's number of electrons = $4.4\,8 \times 10^{-4}$ g

25. Calculate the charge of Avogadro's number of electrons.

Solution

Charge of one electron = $-1.6\,0\,2 \times 10^{23}$C

Charge of $6.0\,2 \times 10^{23}$ electrons = x

$$x = \frac{\text{charge of } 6.02 \times 10^{23} \text{ eletrons} \times (-1.602 \times 10^{-19}\text{C})}{\text{Charge of one eletron}}$$

$$x = -9.6 \times 10^{4}\text{C}$$

Thus the charge of Avogadro's number of electrons is -9.64×10^{4}C

2. Discover of proton

It was discovered by Eugene Goldstein using **anode ray or canal ray experiments**. Canal rays are positively charged ions formed by gases. The charge to mass ratio of hydrogen ion was the highest among all gases. It was also observed that the hydrogen ion has the smallest size among all ionized gases.

Properties Canal rays

- They travel in straight line
- They deflected in electric and magnetic field torward negative plate
- They consisted of material particles and possess energy and momentum.
- They are positively charged

3. Discovery of the Nucleus

Radioactivity:- is the spontaneous emission of particles and or radiation

by radioactive elements. Elements with unstable nuclei that emit radiation / or particles are called **radioactive elements.**

Example:

Uranium (U - 238)
Carbon (C - 14)
Chromium (C - 55)
Radium (Ra - 226) etc

In general, elements whose atomic number is greater than or equal to 83 are radioactive elements.

There are three types of radiations emitted by radioactive elements These are

1. Alpha rays 2. Beta rays 3. Gamma rays

Table 3.4 The Properties of the Types of Rays

Ray	Symbol	Charge	mass in gram	Deflection	Penetration
Alpha	$^{4}_{2}\alpha$ or $^{4}_{2}He^{2+}$	+2	6.65×10^{-24}	To negative plate	Stopped by Paper
Beta	$^{0}_{-1}\beta$ or $^{0}_{-1}e$	-1	9.10×10^{-28}	To Positive plate	Stopped by aluminum
Gamma	$^{0}_{0}\gamma$	0	0	Undeflected	Stopped by lead

Nucleus

It was discovered by E. Rutherford using alpha scattering experiment.

- He allowed alpha particles to pass through gold foil.
- Three main observation were made.

1. Most of the alpha particles went through the thin gold foil with out any change in path. This indicated most of the spaces in an atom was **empty**.
2. A few alpha particles were deflected through small angles. This showed that nucleus has **positive** charge.

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3. Few alpha particles deflected with large angles because of the heavy **nucleus.**

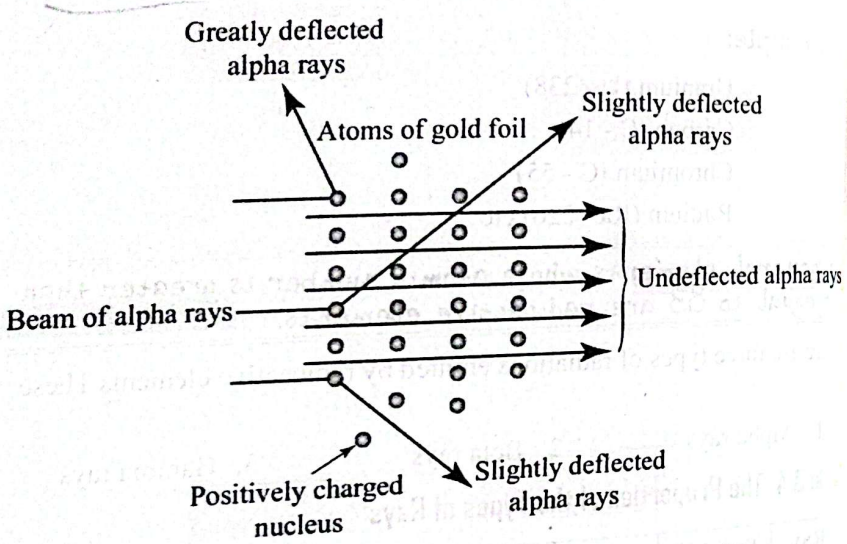

Fig 3.2 The Gold foil Experiment indicated that the Gold Atom Consists of a Central Positively Charged Nucleus

Neutron:-

- Its existence has been predicted by E . Rutherford.
- It was discovered later by the student of E. Rutherford **James Chadwick.**
- It has no charge and found in the nucleus.
- Its mass is similar with the mass of a proton.

Neutron reduces the Proton - proton repulsion and contributes half of the mass of an atom.

Table 3.5 The Three Fundamental Sub - atomic Particles.

Particle	Symbol	Mass in amu	Mass in gram	Charge in coloumb	Charge in unit
Electron	e	0.00055	9.10×10^{-28}	-1.602×10^{-19}	- 1
Proton	p	1.0073	1.673×10^{-24}	$+1.602 \times 10^{-19}$	+ 1
Neutron	n	1.0087	1.675×10^{-24}	0	0

Note:- Proton and neutron are collectively known as **nucleons.** Proton and neutron are found in the nucleus. Electron is found outside the nucleus in atomic shell or energy level.

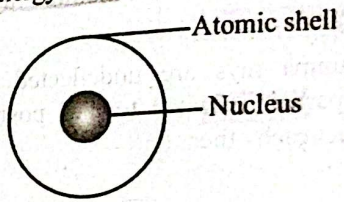

Fig 3.3 Regions of an Atom

Illustrative Examples

26. What did Rutherford's alpha particle experiment show?

A. Electron have negative charge
B. A proton is a hydrogen atom without electron
C. Electrons circle the nucleus an atom in orbits
D. Most of the mass and all of the positive charge of an atom are found in the tiny nucleus.

Solution

Rutherford discovered that an atom has a tiny nucleus with positive charge and heavy mass.

Answer: D

27. Which of the following sub - atomic particles have similar mass?

A. Proton and neutron
B. Proton and electron
C. Neutron and electron
D. Proton, neuron and electron

Answer: A

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28. Which of the following types of rays have the same charge as nuclei of helium atoms?

A. Gamma rays
C. Alpha rays
B. Beta rays
D. X - rays

Solution

Experiments showed that alpha rays were nuclei of helium atoms.

Answer: C

29. Which of the following particles deflected toward positive plate?

A. Neutrons
C. Alpha particles
B. Beta particles
D. Gamma rays

Solution

Neutrons and gamma rays are undeflected because they are chargeless. Beta partial deflected toward positive plate because unlike charge attract each other.

Answer: B

30. Calculate the charge to mass ratio of a proton.

Solution

$m_p = 1.673 \times 10^{-24}$ g

Charge of p = 1.602×10^{-19}C

$$\frac{\text{Charge}}{\text{mass}} = \frac{1.602 \times 10^{-19}\ C}{1.673 \times 10^{-24}\ g} = 9.58 \times 10^{4}\ C/g$$

31. What are the two regions of an atom?

Solution

- Nucleus
- Atomic shell

32. Write the importance of neutron.

Solution

- It reduces the strong electrostatic repulsion forces between protons.
- Half of the mass of an atom is almost the mass of the neutron.

Atomic Number and Mass Number

Atomic number (Z) is the number of protons in an atom.

For a neutral atom : Z = p = e

For charged atom: e = Proton - charge

Mass number (A) is the sum of atomic number and neutron number.

A = Z + n = p + n

n = A - Z

Example:

Aluminium has 13 protons, 13 electrons and 14 neutrons, Determiner its mass number.

A = p + n

A = 13 + 14

A = 27

Mass number is the same as number of nucleons.

Isotopes and Atomic Mass

Isotopes are atoms of the same element with the same atomic number but different mass number.

Representation of Isotopes:- Isotopes can represented by two different ways. These are:

(i) Nuclear Symbolic Representation

$^{A}_{Z}\chi$

A → mass number

χ → symbol

Z → atomic number

Example : $^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$

(ii) Hyphen Notation

χ - A

Example:

C - 12, C - 13, C - 14

Isobars are species with same mass number

Isotones are species with same neutron number

Isoelectronics are species with the same electron number

Atomic mass is the average mass of the atoms of an element. By international agreement, an atom of the carbon isotope called carbon - 12 has exactly 12 atomic mas units (amu).

$$\text{amu} = \frac{\text{mass of one carbon} - 12 \text{ atom}}{12}$$

$$1\text{amu} = 1.66\times10^{-24}\text{ g}$$

Average Atomic mass of an element $= A_1 f_1 + A_2 f_2 + ... A_n f_n$

- $A_1, A_2, ... A_n$ are atomic mass of first second, ... n^{th} isotope respectively
- $f_1, f_2, ... f_n$ are percentage abundance of first, second, ... n^{th} isotope respectively.

Average Atomic mass of an element $= A_1 X_1 + A_2 X_2 + ... A_n X_n$

- $X_1, X_2, ... X_n$ are fraction abundance of first, second, ... n^{th} isotope respectively .

Example:

The atomic mass of the two stable isotopes of boron $^{10}_{5}B$ (19.78%) and $^{11}_{5}B$ (80.22%) are 10.0129 amu and 11.0093 amu respectively. Calculate the average atomic mass of boron?

Solution

$$\% = \frac{1}{100}$$

$$A_1 = 10.0129 \text{ amu}, f_1 = 19.78\% = \frac{19.78}{100}$$

$$A_2 = 11.0093 \text{ amu}, f_2 = 80.22\% = \frac{80.22}{100}$$

Average atomic $= A_1 f_1 + A_2 f_2$

$$= 10.0129 \times \frac{19.78}{100} + 11.0093 \times \frac{80.22}{100} = 10.8 \text{ amu}$$

Illustrative Examples

33. Given the chemical symbol of a certain element as $^{50}_{23}Z$, how many neutrons does it contain?

A. 73 B. 27 C. 23 D. 50

Solution

$$n = A - Z = 50 - 23 = 27$$

Answer: B

34. Isotopes are atoms of the same element that differ in their:

A. Number of proton
B. Number of electron
C. Mass number
D. Atomic number

Solution

Isotopes have **same:**

- Atomic number
- Proton number
- Electron number
- Chemical property

Different:

- Neutron number
- Mass number
- Physical property

Answer: C

35. Gallium has two naturally occurring isotopes. Information about the two isotopes is given in the table below.

Isotope	Atomic mass (amu)	Percentage abundance (%)
Ga - 69	68.93	60.11
Ga - 71	70.92	39.89

What is the average atomic mass of gallium?

A. 69.54 B. 70.29 C. 70.00 D. 69.72

Solution

Average atomic $= A_1 f_1 + A_2 f_2$

$$= 68.93 \times \frac{60.11}{100} + 70.92 \times \frac{39.89}{100} = 69.72 \text{ amu}$$

Answer: D

36. The number of nucleons in $^{238}_{92}U$ is:

A. 92 B. 238 C. 146 D. 330

Solution

Number of nucleons is the same as mass number.

Answer: B

37. Determine the number of electrons, protons and neutrons in:

a. $^{80}_{35}Br$ b. $^{71}_{31}Ga^{3+}$ c. $^{128}_{52}Te^{2-}$

Solution

a. For neutral atom: $Z = p = e^-$

$\Rightarrow Z = p = e = 35$

$n = A - Z = 80 - 35 = 45$

b. For charged atom : $Z = p$

$Z = p = 31$

$e = p - charge = 31 - (+3) = 31 - 3 = 28$

$n = A - Z = 71 - 31 = 40$

c. $Z = p = 52$

$e = p - charge = 52 - (-2) = 52 + 2 = 54$

$n = A - Z = 128 - 52 = 76$

38. Chlorine has two isotopes, Cl - 35 and Cl - 37 with average atomic mass 35.5. Find the percentage abundance of each isotope.

Solution

$A_1 = 35$ $f_1 + f_2 = 100$

$A_2 = 37$ $f_2 = (100 - f_1)\%$

Average atomic mass $= A_1 f_1 + A_2 f_2$

$$35.5 = 35 \times \frac{f_1}{100} + 37\frac{(100 - f_1)}{100}$$

$3550 = 35f_1 + 3700 - 37f_1$

$3550 - 3700 = 35 f_1 - 37 f_1$

$-150 = -2 f_1$

$f_1 = 75\%$

$f_2 = 100 - 75 = 100 - 75 = 25\%$

39. The atomic masses of $^{6}_{3}Li$ and $^{7}_{3}Li$ are 6.0151 amu and 7.0160 amu with fraction abundance of 0.07493 and 0.92507 respectively. Find the average atomic mass of lithium.

Solution

$A_1 = 6.0151$ amu $X_1 = 0.07493$

$A_2 = 7.0160$ amu $X_2 = 0.92507$

Average atomic mass $= A_1 X_1 + A_2 X_2$

$= 6.0.151 \times 0.07493 + 7.0160 \times 0.92507$

$= 6.94$ amu

40. A charged atom has 15 protons, 16 neutrons and 18 electrons. Write nuclear symbolic representation for this atom using a symbol P.

Solution

$A = p + n$ Charge $= P - e$

$= 15 + 16$ $= 16 - 18$

$= 31$ $= -3$

Thus, symbolic representation : $^{31}_{15}P^{3-}$

3.5 Atomic Models

An atomic model is representation of an atom. The ideas about atomic model have been changed many times and we summarize these into five models.

1. The Dalton Model

He thought of atom as solid indivisible sphere. This model is also known as the billiard ball model

Note:- Dalton know nothing about the electrical aspect of the atom

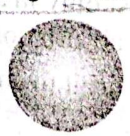


Fig 3.4 The Dalton Atom

2. The Thomson Model

☑ He suggested that in an atom, electrons are scattered throughout a positively charged sphere.

☑ His model is also known as **Plum- pudding model**. This model was discarded as a result of alpha scattering experiment of Rutherford.

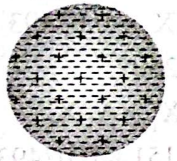


Fig 3.5 The Thomson Atom

3. The Rutherford Model

☑ Based on this model, an atom has a tiny, Positively charged nucleus which electrons whirled at high speed and relatively high distance from it.

☑ This model is also called the **planetary model**. According to this model , the electron that rotates around the nucleus loses energy and fails on the nucleus and the atom collapses. But atoms do not collapse.

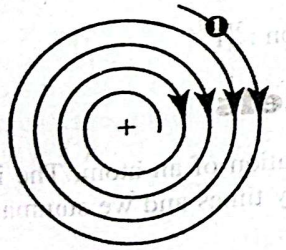


Fig 3.6 Spiral Path of the Electron

4. The Bohr Model

☑ Bohr ammended Rutherford's model in two ways:-

(i) The electron in an atom can exist only in a restricted number of stable orbits in which they neither absorb nor emite energy

(ii) An electron can jump from one orbit to another by **absorbing energy** and falls by emitting energy.

☑ Bohr model introduced energy levels of a hydrogen atom using numbers and letters as

K - shell - 1st energy level
L - shell - 2nd energy level
M - shell - 3rd energy level
N - shell - 4th energy level
O - shell - 5th energy level

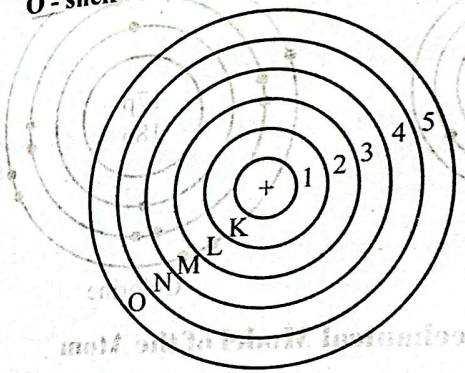


Fig 3.7 The First Five Energy Levels

☑ **Bohr** also suggested that each orbit /energy level/ can accommodate a certain number of electrons.

The maximum number electrons in an orbit can be determined by the formula $2n^2$, where, n = 1, 2, 3, ...

Example:

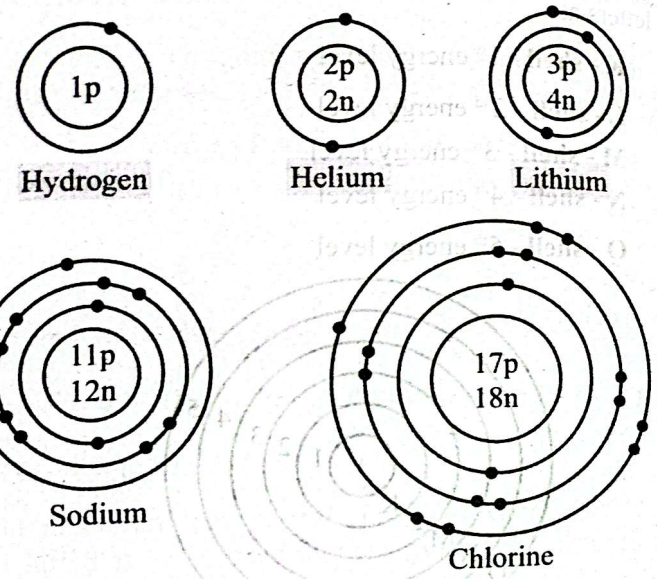
The number of electrons in the first, second and third energy levels are;

- n = 1, number of electrons = $2n^2 = 2\,(1)^2 = 2$ electrons
- n = 2, number of electron = $2n^2 = 2\,(2)^2 = 8$ electrons
- n = 3, number of electron = $2n^2 = 2\,(3)^2 = 18$ electrons

But, regardless of its maximum capacity, any outer most energy level of an atom cannot hold more than eight electrons.

Bohr's diagrammatic representation using energy levels /shells/ orbits in

some atoms.



5. The Quantum Mechanical Model of the Atom

- It is also called **shr "O" ding** or wave mechanical model.
- It describes electron to be in constant motion some where outside the nucleus.
- The region in space around the nucleus where there is high probability of getting electron is called **orbital**.
- This model introduced sub-shells designated by small letters s,p,d,f,g,..

Number of orbitals = n^2

Number of sub - shells = n

Maximum number of electron in an orbital = 2

Maximum number of electron in main shell = $2n^2$

Table 3.6 Number of Orbitals and Maximum Electrons in Sub-shell

Sub - shell	Number of orbital	Number of electrons
s	1	2
p	3	6
d	5	10
f	7	14

Table 3.7 Number of Sub-shell and Electrons in the first five Energy Levels.

Energy level	Symbol	Number of Sub - shell	Types of Sub-shells	Maximum number of electrons in main shell
n = 1	K	1	s	2
n = 2	L	2	s, p	8
n = 3	M	3	s, p, d	18
n = 4	N	4	s, p, d, f	32
n = 5	O	5	s, p, d, f, g	50

Illustrative Examples

41. Whose model was discarded as a result of Rutherford's model?

A. Dalton's Model
B. Thomson's model
C. Bohr's model
D. Quantum model

Solution

Electron and proton have opposite charges and cannot be found together. This idea was explained by Rutherford.

Answer: B

42. Which model of the atom can be pictured as "Planetary atom"?

A. Bohr's model
B. Rutherford's model
C. Thomson's model
D. Dalton's model

Answer: B

43. The atomic model of Thomson, Dalton, Rutherford and Bohr, all

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support which of the following statements?

A. All matter is composed of atoms
B. Atoms are mostly composed of empty space
C. Atoms are composed of electrons, neutrons and protons.
D. Electron orbit the nucleus

Solution

According to all atomic models, matter is composed of atoms.

Answer: A

44. What is the maximum number of electrons in the N-shell?

A. 2 B. 8 C. 18 D. 32

Solution

N-shell is the fourth shell (n = 4), number of electrons = $2n^2 = 2(4)^2 = 32$

Answer: D

45. Which of the following atomic models was developed with a proposition that electrons can be described only interns of the probability of their location?

A. Rutherford's model C. Quantum mechanical model
B. Dalton's model D. Bohr's model

Answer: C

46. A particular shaped Volume of space in an atom where the probability of finding an electron is maximum is known as:

A. An orbital C. A sub- energy level
B. An orbit D. A shell

Answer: A

47. Determinerthenumberofelectrons,orbitalsandsub-shellsinthefifthshell

Solution

Number of electrons = $2n^2 = 2 \times 3^2 = 18$

Number of orbitals = $n^2 = 3^2 = 9$

Number sub - shell = n = 3 (s, p, d)

48. Why did scientists consider Rutherford's nuclear model of an atom incomplete?

Solution

Because his model lacked details on how electrons of an atom are arranged around the nucleus.

49. Explain the relationship between principal quantum number (n) and

a. Energy of orbitals b. Size of orbitals

Solutions

As the value of n increases, energy and size of orbitals also increase.

50. Arrange the first four energy levels in increasing order of energy?

Solution

K < L < M < N

51. In quantum mechanical model concept, each energy level is described by a number called _____

Answer: Principal quantum number.

52. Arrange the five atomic model in order of historical development that led to the modern model of the atom?

Solution

Dalton's Model → Thomson model → Rutherford's model → Bohr model → Quantum mechanical model

3.6 Electron Configuration

It is the arrangement of electrons about the nucleus of an atom according to quantum mechanical model description. There are different types of electron configurations.

(i) Main Energy Level Electron Configuration

- Main energy levels /orbits/ were introduced by the student of Rutherford, Bohr.
- They are atomic shells in which the electron is found relative to atom's nucleus.
- Energy level with lower energy (smaller value of n) filled first.

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Table 3.8 The Main Energy level Electron Configuration of the First 20 Elements.

Element	Symbol	Atomic number	Electron Con figuration in main energy level				
			K	L	M	N	O
Hydrogen	H	1	1				
Helium	He	2	2				
Lithium	Li	3	2	1			
Beryllium	Be	4	2	2			
Boron	B	5	2	3			
Carbon	C	6	2	4			
Nitrogen	N	7	2	5			
Oxygen	O	8	2	6			
Fluorine	F	9	2	7			
Neon	Ne	10	2	8			
Sodium	Na	11	2	8	1		
Magnesium	Mg	12	2	8	2		
Aluminium	Al	13	2	8	3		
Silicon	Si	14	2	8	4		
Phosphorus	P	15	2	8	5		
Sulphur	S	16	2	8	6		
Chlorine	Cl	17	2	8	7		
Argon	Ar	18	2	8	8		
Potassium	K	19	2	8	8	1	
Calcium	Ca	20	2	8	8	2	

Note:

- The electrons that are found in the valence shell or outmost shell are called **valence electrons**
- The minimum and maximum number of valence electrons is **one and eight** respectively.

Example:

Determine the valence electrons in an element whose atomic number is 17

Solution

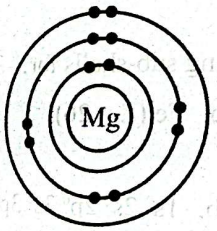
- Electron configuration: 2, 8, 7
- Valence shell is the 3rd shell
- Valence electrons = 7

(ii) Electron Configuration using Energy Diagrams

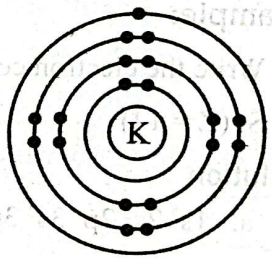
This type of electron configuration also follows the same principle as the above but using circular paths (orbits) :

Example:

a. Mg (Z = 12): 2, 8, 2



b. K(Z = 19): 2, 8, 8, 1



(iii) Electron Configuration Using Sub-shells

- Sub - shells with the lowest energy filled first. This principle is called **the Aufbau or building up** principle.

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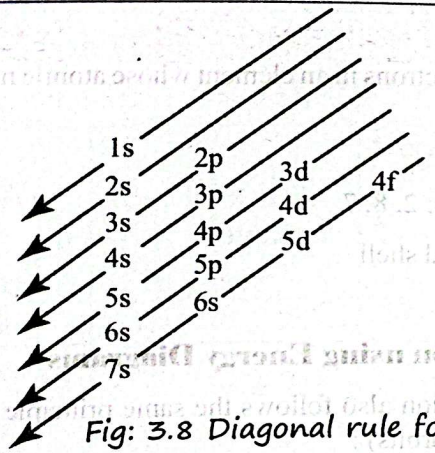


Fig: 3.8 Diagonal rule for Filling Electrons

If you list the sub-shells following the arrows, you get electron configuration of most atoms.

Example:

Write the electron configuration using sub-shells for:

a. Si (Z = 14)

b. Fe (Z = 26)

Solution

a. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

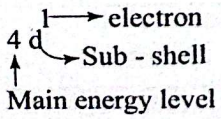
b. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

Note : Exception to Building up principle

$_{24}Cr$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$

$_{29}Cu$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$

Species that have the same electron configuration called isoelectronic species



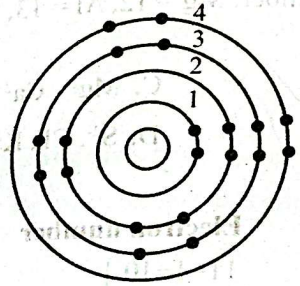
Illustrative Examples

53. Which of the following is not a valid electron configuration?

A. 2, 8, 8, 2 B. 2, 8, 9, 1 C. 2, 6 D. 2, 8, 4

Answer: B

54. The following diagram shows Bohr model for a certain atom.



In which energy level (s) do the valence electrons located?

A. 2 and 3 B. 1 C. 3 and 4 D. 4

Solution

The valence electrons are found in the outer most shell i.e in the 4^{th} shell.

Answer: D

55. How many electrons enter in to 3d sub - shell of an atom whose atomic number is 21?

A. 9 B. 1 C. 0 D. 3

Solution

Electron configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$

➥ There is one electron in 3d sub-shell.

Answer: B

56. In the notation $3d^1$ for the electron configuration, 3, d and 1 respectively stands for:

A. Main energy level, sub-shell and electron

B. Electron, sub-shall and main energy level.

C. Main energy level, electron and sub-shall

D. Sub - shall, main energy level and electron.

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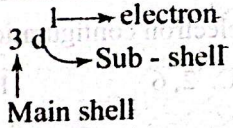
Solution



Answer: A

57. Which group Contains chemical species with the same electron configuration? (Atomic number: Mg = 12, Al = 13, S = 16, Cl = 17, Ca = 20, K = 19 Na = 11)

A. Mg, Ca, Na
B. K^+, Al^{3+}, S^{2-}
C. Mg^{2+}, Ca^{2+}, Na^+
D. S^{2-}, Cl^-, K^+

Solution

Ion	Electron number	
Na^+	11−1=10	ioelectronic
Mg^{2+}	12−2=10	
Al^{3+}	13−3=10	
S^{2-}	16 − (−2) = 18	isoelectronic
Cl^-	17 − (−1) = 18	
K^+	19 − 1 = 18	
Ca^{2+}	20 − 2 = 18	

Answer: D

58. Write electron configuration using main energy levels for:

a. $_{20}Ca^{2+}$ b. $_{31}Ga$ c. $_{3}N^{3-}$ d. $_{38}Sr$

Solution

a. e = 20 -2 = 18
Electron configuration: 2, 8, 8.

b. e = 31: 2, 8, 18, 3
c. e = 7- (-3) = 10: 2, 8
d. e = 38: 2, 8, 18, 8, 2

59. The main energy with the lowest energy is _____

Solution

The K - shell (the first energy level).

60. Write the electron configuration using sub-shell for

a. $_{15}P^{3-}$ b. $_{28}Ni$

Solution

a. $e = 15 - (-3) = 18 : 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

b. $e = 28 : 1s^2\ 2s^2\ 2p^6\ 3s^1\ 3p^6\ 4s^1\ 3d^8$

Solved Problems

61. In which countries did the atomic theory originate as philosophical concept?

Solution

In ancient India and Greece

62. Who discovered an electron and determined the charge to mass ratio of it?

Solution

J . J Thomson

63. Atoms of the same element have the same mass was found to be wrong Why?

Solution

Because of the discovery of isotopes.

64. Explain why atoms of different elements may not be different in all aspects?

Solution

➥ Atoms of different element may have the same mass number

Example:

$^{14}_{6}C$, $^{14}_{7}N$

Atoms of different element may have the same number of neutrons

Example:

$^{23}_{11}Na$ and $^{24}_{12}Mg$

65. In the discovery of electrons, J. J Thomson conducted many experiments on cathode rays. List them.

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Solution

- He was studied the straight line travel of cathode rays in the discharge tube by placing a small object between the cathode and anode.
- He placed a light paddle wheel between cathode and to study the particulate nature of cathode rays.
- He passed the cathode rays through magnetic and electric fields to investigate the negative charges of cathode rays.

66. Compare the mass of proton and electron.

Solution

$m_e = 9.10 \times 10^{-28}$g = mass of an electron

$m_p = 1.673 \times 10^{-24}$g = mass of a proton

$$\frac{m_p}{m_e} = \frac{1.673 \times 10^{-24}\,g}{9.10 \times 10^{-28}\,g} = 1838$$

$\Rightarrow m_p = 1838\ m_e^-$

67. Write the names of the three isotopes of hydrogen

Solution

Protium $\left({}_1^1H\right)$

Deuterium $\left({}_1^2H\right)$

Tritium $\left({}_1^3H\right)$

68. Carbon has three isotopes with neutron number 6, 7 and 8. Its atomic number is 6. Represent the three isotopes of carbon using nuclear and hyphen notations.

Solution

$A_1 = Z + n_1 = 6 + 6 = 12$

$A_2 = Z + n_2 = 6 + 7 = 13$

$A_3 = Z + n_3 = 6 + 8 = 14$

Nuclear symbolic representation ${}_6^{12}C$, ${}_6^{13}C$, ${}_6^{14}C$

Hyphen notation: C - 12, C - 13, C - 14

69. An element has two isotopes A - 16 and A -18, The average atomic mass of A is 16.5. Calculate the percentage and fraction abundance of each isotope.

Solution

$A_1 = 16$ $\qquad f_1 + f_2 = 100\%$

$A_2 = 18$ $\qquad f_2 = (100 - f_1)\,\%$

Average atomic mass = 16.5

Average atomic mass $= A_1 f_1 + A_2 f_2$

$$16.5 = 16 \times \frac{f_1}{100} + 18\frac{(100 - f_1)}{100}$$

$1650 = 16f_1 + 1800 - 18f_1$

$f_1 = 75\%$ and $f_2 = 25\%$

$X_1 = 0.75$ and $X_2 = 0.25$

70. Is there an atom with more than one protons but no neutron?

Solution

No, because neutrons reduce the proton-proton electrostatic force of repulpsion

71. Natural copper is a mixture of only two isotopes one of which has 29 protons, and a relative abundance of 69.17 %. The relative atomic mass of natural copper is 63.6. How many neutrons must be there in the second isotope?

Solution

$A_1 = p + n = 29 + 34 = 63$

$f_1 = 69.17\%$

$f_2 = 100 - f_1 = 100 - 69.17 = 30.83\%$

$A_2 = ?$

Av atomic mass $= A_1 f_1 + A_2 f_2$

$$63.6 = 63 \times \frac{69.17}{100} + A_2 \times \frac{30.83}{100}$$

Solving for $A_2 = 65$

$A_2 = p + n_2$

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$n_2 = A_2 - p = 65 - 29 = 36$

Thus, the number of neutrons in the 2nd isotope is 36.

72. An arbitrary element "X" has two naturally occuring isotopes.

Isotope	Atomic mass in amu	Percentage abundance (%)
X - 40	40	80
X - 44	44	20

Solution

Solve by your own!

Answer: 40.80amu

73. Carbon has two isotopes, C - 12 and C - 13 that occurs in 9 : 1 abundance ratio. Determine the relative atomic mass of carbon.

Solution

Convert the ratio in to abundance fraction.

$$X_1 = \frac{9}{1+9} = 0.9$$

$$X_2 = \frac{1}{1+9} = 0.1$$

$$\text{Av. atomic mass} = A_1 X_1 + A_2 X_2$$

$$= 12 \times 0.9 + 13 \times 0.1 = 12.1\text{amu}$$

74. Two Particles A and B have the following composition.

A. 15 protons, 16 neutrons, 15 electrons

B. 15 Protons, 16 neutrons, 18 electrons

a. What is the relationship between these Particles?

b. Will these particles have similar chemical properties?

Solution

a. B is anion of A

b. No, because they have different electron number

75. Identify the following subatomic particles.

a. The number of these in the nucleus is equal to atomic number.

b. The particles that are gained or lost when ions are formed.

c. The particles that are not found in the nucleus.

d. The particles that have lower mass than other particles.

e. The particle that has no charge.

Solution

a. Protons

b. Electrons

c. Electrons

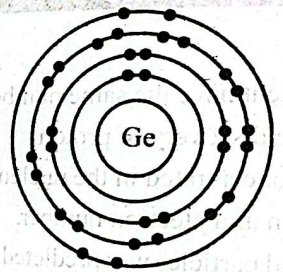
d. Electrons

e. Neutrons

76. Draw Bohr diagram for germanium $\left({}^{73}_{32}\text{Ge}\right)$.

Solution

2, 8, 18, 4



77. Explain the term model.

Solution

Model is the mind's picture of what areal thing would look like. Picture about things which cannot be easily seen directly.

78. An atom has electron configuration of $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2 3d^{10}\ 4p^4$. If its mass number is 128, What will be the number protons and neutron?

Solution

$$p = e = 2 + 2 + 6 + 2 + 6 + 2 + 10 + 4 = 34$$

$$n = A - p = 128 - 34 = 94$$

79. Write the electron configuration of the following elements using sub - shells.

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a. Cr (Z = 24) b. Cu (Z = 29)

Solution

a. Exceptionally, these two elements have different electron configurations.

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$

b. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$

80. The valence shell configuration of an element is $4s^2\ 4p^5$. Determine its atomic number.

Solution

First write the complete electron configuration.

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$

➥ $e = Z = 2 + 2 + 6 + 2 + 6 + 2 + 10 + 5 = 35$

Supplementary Exercise

I. True/ False Items

1. Different atoms of the same element have the same number of protons.
2. Canal rays experiment lead to the discovery of proton.
3. Most of the mass of an atom is concentrated in the nucleus.
4. Mass number is the sum of proton and electron number.
5. The presence of positively charged particles was predicted by Goldstein.
6. According to Dalton, an atom is divisible.
7. The word "atom" comes from the Latin word atomos meaning indivisible.
8. According to Bohr model, electrons couldn't spiral in to the nucleus.
9. An atom is neutral because it has equal number of protons and neutrons.
10. Some of the elements have fractional atomic masses because of existence of isotopes.

II. Matching Items

A	B
11. Thomson's Model	A. Planetary Model
12. Dalton's Model	B. Billard ball Model
13. Rutherford's Model	C. Plum - Pudding Model
14. Bohr's Model	D. Orbital
15. Quantum Mechanical model	E. Orbit

III. Multiple Choice Items

16. A sample of water taken from ice, rain or sea has the same ratio by mass of hydrogen and oxygen. This chemical fact explains the law of:

A. Definite proportion C. Conservation of mass

B. Multiple proportion D. Mass of action

17. The discovery of small but massive positively charged nucleus in the atom led to which model of the atom?

A. Quantum mechanical model C. Plum - pudding model

B. Bohr model D. Planetary model

18. How many neutrons are there in phosphorus atom having a mass number of 31 and atomic number 15?

A. 31 B. 46 C. 15 D. 16

19. A model of an atom that describes an atom as a positive sphere where electrons are embedded in a sphere is:

A. Thomson's Model C. Quantum mechanical Model

B. Dalton's Model D. Rutherford's Model

20. Given two hypothetical compounds A and B that contain sulphur and oxygen according to the information give in the table below

Compound	Mass of sulphur in gram	Mass of oxygen in grams
A	50	50
B	50	75

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What is the ratio of masses of oxygen in these compounds?

A. 1 : 1 B. 2 : 3 C. 2 : 5 D. 1 : 2

21. The actual or absolute mass of an electron is:

A. 9.1×10^{-31}kg C. 1.60×10^{19}g

B. 9.1×10^{-28}kg D. 1.76×10^{8}g

22. An electron has:

A. a charge of + 1 and a smaller mass than a proton

B. a charge of + 1 and the same mass a proton

C. a charge of - 1 and a smaller mass than a proton

D. a charge of - 1 and the same mass a proton

23. Which of the following are usually found in the nucleus of an atom?

A. Proton and electron only C. Proton, neutron and electron

B. Proton and neutron only D. Neutron only

24. An atom has an atomic number of 31 and a mass number of 70. How many electrons will it have in its valence shall?

A. 5 B. 4 C. 2 D. 3

25. Which one of the following is the smallest particle of a substance that takes part in a chemical reaction?

A. An atom B. Anion C. A formula unit D. A molecule

26. An ion that consists of 7 protons, 7 neutrons and 10 electrons has a net charge of

A. + 3 B. + 4 C. - 3 D. - 4

27. The number of protons, electrons and neutrons in $_{30}{}^{65}Zn^{2-}$ respectively are:

A. 30, 30, 35 B. 30, 28, 35 C. 30, 28, 65 D. 30, 30, 65

28. The charge of an electron was determined by:

A. E. Rutherford C. J. Dalton

B. J. J. Thomson D. R. Millikan

29. A mono - atomic ion that has 13 protons and a +3 charge:

A. has 10 neutrons C. has a symbol Ne

B. has 16 electrons D. is isolectronic with Ne

30. The elements $^{14}_{6}C$ and $^{14}_{7}N$ have the same

A. Chemical property C. Proton number

B. Mass number D. Neutron number

31. What is the total number of sub - atomic particles in $^{108}_{47}Ag$?

A. 4 B. 108 C. 155 D. 61

32. The maximum number of electrons in the second shell is:

A. 4 B. 8 C. 2 D. 16

33. $^{39}_{19}K$ and $^{40}_{20}Ca$ have the same number of

A. Nucleons B. Protons C. Electrons D. Neutrons

34. Which sub - shall is found in all main energy levels?

A. s B. P C. d D. f

35. The aufbau principle is one that governs

A. Vapour pressure C. Electronic configuration

B. Proton distribution D. Atomic models

36. An atom is electrically neutral because it contains equal number of

A. Proton and electron C. Neutron and electron

B. Proton and neutron D. Nucleon and electron

37. The atoms $^{40}_{20}A$ and $^{41}_{x}B$ are isotopes of the same element .The value of x in the representation of B is:

A. 2 B. 10 C. 15 D. 20

38. A Model in chemistry is used in the following

A. Giving a mind's picture of an idea.

B. Predicting two possible ways a reaction can take place.

C. Showing the mass of an atom.

D. Experimenting for the first time.

39. The relative atomic mass of neon is 20.2. It consists of the isotopes $^{20}_{10}Ne$ and $^{22}_{10}Ne$. $^{20}_{10}Ne$ has a percentage abundance in naturally occurring neon of:

A. 80 B. 90 C. 0.9 D. 50

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40. The atomic number of an atom is always equal to:
   A. Proton number
   B. Mass number
   C. Neutron number
   D. Electron number

41. Which of the following statements best describes Bohr's postulate about the electronic structures of an atom?
   A. Electrons occupy restricted orbits and their transition between the orbits is accompanied by their absorption or emission of energy.
   B. Electrons occupy random space around the nucleus and transition is not allowed for the electrons under any circumstance
   C. When the electrons fall from higher (excited) state to the lower (ground) state, they absorb energy.
   D. When electrons jump from a lower to higher energy state, they emit a fixed amount of energy.

42. According to the quantum mechanical model of the atom.
   A. Electron are embeded in an atom and remain stationary.
   B. The atom is a positively charged solid sphere
   C. Orbitals are represented as clouds of charge
   D. Electrons can revolve in all orbits with all possible radii.

43. Which sequence represents a correct order of the historical development that led to the modern model of the atom?
   A. Most of the atom is empty → the atom is a hard sphere → electrons exist in orbitals outside the nucleus.
   B. The atom is a hard sphere → most of the atom is empty → electrons exist in orbitals outside the nucleus.
   C. Most of the atom is empty → electrons exist outside the nucleus → the atom is a hard sphere.
   D. The atom is a hard sphere → electrons in orbitals outside the nucleus → most of the atom is empty.

44. The number of sub - shells in the fifth shell is:
   A. 25 B. 10 C. 5 D. 50

45. The electron configuration for copper $(_{29}Cu)$ is :
   A. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^9$
   B. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$
   C. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$
   D. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 4d^{10}$

IV. Short Answer Item

46. Determine the number of electrons, protons and neutrons.
   a. $^{197}_{78}Pt$
   b. $^{97}_{42}Mo^{2+}$
   c. $^{75}_{33}As^{3-}$

47. Write electron configuration by using sub - shells.
   a. Br (Z = 35)
   b. Cd (Z = 48)

48. Write electron configuration by using main energy levels for
   a. S^{2-} (Z = 16)
   b. Si^{4+} (Z = 14)

49. The two bromine isotopes ^{80}Br and ^{81}Br occurs in 1 : 1 ratio calculate the average atomic mass of bromine.

50. Copper has two naturally occuring isotopes.

Isotope	% abundance	Atomic mass in amu	Average atomic mass
1	69.150	x	63.546
2	30.850	$x + 2$	

What is the value of x ?

Answer to Suplementary Exercise

I. 1. True 2. True 3. True 4. False 5. True 6. False
7. False 8. True 9. False 10. True

II. 11.C 12. B 13. A 14. E 15. D

III. 16. A 17. D 18. D 19. A 20.B 21. A
22.C 28. D 34. A 40. A
23.B 29. D 35. C 41. A
24.D 30. B 36. A 42. C
25.A 31. C 37. D 43. B
26.C 32. B 38. A 44. C
27.B 33. D 39. B 45. B

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46. a. e = 78, p = 78, n = 119 c. e = 36, p = 33, n = 42

b. e = 40, p = 42, n = 55

47. a. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$

b. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}$

48. a. 2, 8, 8 b. 2, 8

49. 80.5 amu

50. 62.92g

UNIT 4

Periodic Classification of Elements

INTRODUCTION

A systematic classification of elements allows us to grasp maximum possible information with short time, low cost and least effort. The importance of periodic classification of elements lies in the fact that an element which is a member of a certain group will show the major characteristic of the family without losing its unique properties.

4.1 Early Attempts at Classification Dobereiner and Newland

A. Dobereiner Triad

- **Dobereiner** a German chemist, in 1817 discovered groups of three related elements which he termed a **Triad**
- He found three elements with similar properties and the atomic mass of the middle element was about midway between the other two elements.

Table 4.1 Some of Dobereiner's Triad

Name	Atomic Mass	Name	Atomic Mass	Name	Atomic Mass
Lithium	7	Calcium	40	Chlorine	35.5
Potassium	39	Barium	137	Iodine	127
Average	23	Average	88.5	Average	81.3
Sodium	23	Strontium	87.6	Bromine	80

Limitation of Law of Triad

The law of triad was not applied even for all the elements discovered at that time.

B. Newland's Law of Octaves

This law was proposed to explain the property similarities which occurred with every eighth elements when they were arranged in increasing order of atomic mass

Table 4.2 Newlands Law of Octaves, Arrangement of Elements.

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
K	Ca					

For Example:

Based on the law of octaves:

Li and Na
B and Al } have similar properties.
N and P

Limitations of the Law Octaves

- This law does not work after calcium.
- It failed when noble gases were discovered because they became the ninth elements.

Illustrative Examples

1. The earliest of the attempts to classify elements was that of

A) Mendeleev's periodic classification
B) Newland's law of octaves
C) Dobereiner's triad
D) Meyer's classification

Solution

Dobereiner's triad was the earliest attempt to classify elements.

Answer: C

2. Which of the following arrangements of elements cannot be considered as a triad?

A) $^{7}Li, ^{23}Na, ^{39}K$
B) $^{9}Be, ^{24}Mg, ^{40}Ca$
C) $^{32}S, ^{79}Se, ^{128}Te$
D) $^{19}F, ^{35.5}Cl, ^{80}Br$

Solution

If the atomic mass of the middle element is nearly equal to the average of the other two, the given arragment will be considered as a triad.

A) $23 = \frac{7+39}{2} \Rightarrow 23 = 23$ Thus, it is a triad

B) $24 = \frac{9+40}{2} \Rightarrow 24 \approx 24.5$, it is a triad

C) $79 = \frac{32+128}{2} \Rightarrow 79 \approx 80$, it is a triad

D) $35.5 = \frac{19+80}{2} \Rightarrow 35.5 \neq 49.5$, it is not a triad

Answer: D

3. Based on Newland's law of octaves table, which pair of elements have similar properties?

A) Li and F B) O and S C) Na and Ca D) N and Si

Solution

Refer the table every eighth element has similar property with the first.

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Answer: B

4. What is the similarity between Dobereiner's triad and Newland's law of octaves?

Solution

Both arranged elements based on atomic mass.

5. Write two triads by refering the modern periodic table.

Solution

- ^{31}P, ^{75}As, ^{121}Se
- ^{39}K, ^{85}Rb, ^{133}Cs

6. Explain the contribution of Newland's law of octaves to the development of modern periodic table

Solution

Newland proposed for the first time the properties elements.

4.2 Mendeleev's Periodic Table

- Demitri Medeleev, a Russian chemist and the German chemist **Lothar Meyer**, each proposed the periodic law.
- They proposed the periodic law that states " the properties of elements is the periodic function of their atomic mass"

Mendeleev:

- Left blanks in his table for undiscovered elements.
- Predicted properties and masses of unknown elements correctly.
- Organized his table in columns and rows (periods)
- His table consisted of only 63 elements
- These elements were arranged in 7 periods and 8 groups.
- His periodic table was later modified due to the discovery of noble gases.

Table 4.3 Mendeleev's Predictions and Actual Properties

Ekasilicon (Es) Predicted Properties	Germenium (Ge) Actual Properties
➥ Atomic mass = 72	➥ Atomic mass = 72.60
➥ High melting point	➥ Melting point = 958°c
➥ Density = 5.5g/cm³	➥ Density = 5.36 g/cm³
➥ Dark gray metal	➥ Gray metal
➥ Will obtain from K_2 Es F_6	➥ Obtain from K_2 Ge F_6
➥ Slightly dissolved by HCl	➥ Not dissolved by HCl
➥ Will from Es O_2	➥ Forms oxides (Ge O_2)
➥ Density of Es O_2 = 4.7 g/cm³	➥ Density of Ge O_2 = 4.70 g/cm³

Advantage of Mendeleev's Periodic Table

- Leaving gaps or blank spaces for undiscovered elements initiated other scientists to discover them.

Drawbacks of Mendeleev's Periodic Table

- No positions for isotopes
- Wrong order of the atomic masses of some elements

Lother Meyer

- Plotted the atomic volume versus atomic mass curve.
- Elements in the same position of the curve have similar properties.
- The atomic volume of the elements changes periodically, initially decreases then increases.
- The atomic volume increases in a group.

$$\text{Atomic Volume} = \frac{\text{Gram atomic mass of an element}}{\text{Density in gram per milliliter}}$$

Illustrative Examples

7. Mendeleev and Lother Meyer arranged elements based on increasing order of their ____

A) Atomic number
B) Atomic mass
C) Electron configuration
D) Neutron number

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Answer: B

8. All of the following are true about Mendeleev's periodic table. EXCEPT
   A) It consisted of only 63 elements.
   B) Some vacant spaces were specified for undiscovered elements
   C) The elements are arranged in 7 periods and 8 groups
   D) Noble gases were included in his periodic table.

Solution

Noble (inert) gases were not included in Mendeleev's table since they were not discovered at that time.

Answer: D

9. Who plotted the atomic volume Vs atomic mass curve to study the properties of elements?
   A) Newlands
   B) Dobereiner
   C) L. Meyer
   D) D. Mendeleev

Answer: C

10. List some elements that have similar properties based Lother Meyer's curve.

Solution

- **Alkali metals** (Li, Na, K, Rb, Cs)
- **Alkaline earth metals** (Be, Mg, Ca, Sr, Ba)
- **Halogens** (F, Cl, Br, I)

11. More attension was given to Mendeleev's work than Lother Meyer. why?

Explanation

Mendeleev left gap for undiscovered elements and predicted their properties. This initiated other scientists to discover them. That is why a better emphasis was given to him.

12. State Mendeleev's periodic law.

Solution

The physical and chemical properties of elements are the periodic function of their atomic masses.

4.3 Modern Periodic Law

- The drawbacks of Mendeleev's periodic table led **Henry Moseley** to study other property of elements using **X - ray** experiment.
- **Moseley's X - ray** experiment showed that the nucleus of each element has integral positive charge, **the atomic number (Z)**
- **Moseley's law:** $\sqrt{\nu} \propto Z$, ν - frequency of X - ray.
- As a result of Moseley's work, periodic law was modified to **Modern periodic law.**
- Modern periodic law states that the properties of elements is a periodic function of their **atomic number.**

	1	2											13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg	3	4	5	6	7	8	9	10	11	12	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	*	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Ti	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	**	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fi	115 Mc	116 Lv	117 Ts	118 Og

Lanthanides *	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides **	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Modern Periodic Table

- It contains groups (vertical columns) and periods (horizontal rows)
- It has 18 vertical columns but 16 groups because VIIIB composes three vertical colums
- **A group** is also called a family because elements in the given group have similar chemical properties
- The 16 - groups are classified into 8 'A' groups (IA - VIIIA) and 8 'B' groups (IB - VIIIB)

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Table 4.4 Common Names of 'A' Group Elements and their Valence Shell Configuration

Group	Common name	Valence shell configuration
IA	Alkali metals (Except hydrogen)	ns^1
IIA	Alkaline earth metals	ns^2
IIIA	Boron family	$ns^2\ np^1$
IVA	Carbon family	$ns^2\ np^2$
VA	Nitrogen family	$ns^2\ np^3$
VIA	Oxygen family	$ns^2\ np^4$
VIIA	Halogen	$ns^2\ np^5$
VIIIA	'O' group, noble or inert gases	$ns^2\ np^6$ (except He : $1s^2$)

Table 4.5 The Seven Periods, the Number of Elements they Contain and Orbitals Occupied

Period	Number of elements	Orbitals occupied
1	2 ($_1$H and $_2$He)	1s
2	8 ($_3$Li - $_{10}$Ne)	2s, 2p
3	8 ($_{11}$Na - $_{18}$Ar)	3s, 3p
4	18 ($_{19}$K - $_{36}$Kr)	4s, 3d, 4p
5	18 ($_{37}$Rb - $_{54}$Xe)	5s, 4d, 5p
6	32 ($_{55}$Cs - $_{86}$Rn)	6s, 4f, 5d, 6p
7	Incomplete $_{87}$Fr - ?	7s, 5f, 6d, 7p

☑ Elements in the same period have:
- ➥ Same number of shells
- ➥ Consecutive atomic number

☑ Period 1,2 and 3 short periods

☑ Periods 4 and 5 are long periods

☑ Period 6 is the longest period.

Number of shells = Period

Valance electron = Group

Example Determine the group and period of Mg (Z = 12)

Solution

2, 8, 2 or $1s^2\ 2s^2\ 2p^6\ 3s^2$, Valence electrons = 2

$\Rightarrow$ Group = IIA, valence shell = 3

$\Rightarrow$ Period = 3

Valence shell configuration	Group
ns^1	IA
ns^2	IIA
$ns^2\ np^1$	IIIA
$ns^2\ np^2$	IVA
$ns^2\ np^3$	VA
$ns^2\ np^4$	VIA
$ns^2\ np^5$	VIIA
$ns^2\ np^6$	VIIIA

For 'B' Group Elements, Use the Following Simple Rule or Method

Valence shall Configuration	Group
$ns^1\ (n-1)\ d^{10}$	IB (only s - electrons)
$ns^2\ (n-1)\ d^{10}$	IIB (only s - electrons)
$ns^2\ (n-1)\ d^1$	2 + 1 = IIIB (s + d electrons)
$ns^2\ (n-1)\ d^2$	2 + 2 = IVB (s + d electrons)
$ns^2\ (n-1)\ d^3$	2 + 3 = VB (s + d electrons)
$ns^1\ (n-1)\ d^5$	1 + 5 = VIB (s + d electrons)
$ns^2\ (n-1)\ d^5$	2 + 5 = VIIB (s + d electrons)
$ns^2\ (n-1)\ d^6$ $ns^2\ (n-1)\ d^7$ $ns^2\ (n-1)\ d^8$	VIIIB

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Illustrative Examples

13. According to the modern periodic law

A) The properties of elements is the periodic function of their atomic masses

B) No two elements have the same atomic mass

C) No two elements can have the same mass number

D) The properties of the elements are periodic functions of their atomic number

Solution

Based on modern periodic law:

- Elements are arranged based on their atomic number or electron configuration
- No two elements can have the same atomic number

Answer: D

14. How many elements are there in the fifth period?

A) 8 B) 2 C) 18 D) 32

Answer: C

15. How many periods and groups are there in the modern periodic table respectively?

A) 7, 16 B) 7, 18 C) 16, 7 D) 18, 7

Solution

There are 7 periods and 16 groups but 18 vertical columns in the modern periodic table.

Answer A

16. To which group and period of the periodic table does an element with atomic number 32 belong?

A) Group IIB and period 2

B) Group IIA and period 2

C) Group IVA and period 4

D) Group IIIA and period 4

Solution

$Z = 32 : 2, 8, 18, 4$ or $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^2$

Valence electrons = 4 ($4s^2\ 4p^2 = 2 + 2 = 4$)

$\Rightarrow$ Group = IVA

Valence shall = 4th shall

$\Rightarrow$ Period = 4

Answer: C

17. An element has the valence shell electron configuration of $2s^2\ 2p^3$. To which group and period does it belong?

A) IIA and 5 B) VA and 2 C) IVA and 4 D) IIIA and 2

Solution

Valence shell = 2 $\Rightarrow$ period = 2

Valence electrons = 2 + 3 = 5, group = VA

Answer: B

18. Determine the group and period of an element with atomic number 25.

Solution

Electron configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$

$\Rightarrow$ It is 'B' group electron because the last electron is in d - sub shall

Valence shell configuration: $4s^2\ 3d^5$

$\Rightarrow$ Group = 2 + 5 = 7 (VIIB)

Valence shell = 4th shell

$\Rightarrow$ Period = 4

19. The electron configuration of an element that is found in group IIA and period 4 is ____

Solution

Group = IIA $\Rightarrow$ Valence electrons = 2

Period = 4 $\Rightarrow$ Valence shell = 4th shell

$\therefore$ Electron configuration : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$ or 2, 8, 8, 2

20. Which group in the modern periodic table contains elements in the three states of matter at room temperature?

Solution

Group VIIA (Halogens)

- Chlorine and fluorine are gases

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- Bromine is liquid
- Iodine and astatine are solids

21. Why group IA elements are commonly known as alkali metals?

Explanation

Because they react with water and form soluble bases called alkali

22. The shortest and longest periods in the modern periodic table respectively are _____ and _____

Answer: Period 1 and Period 6

Classification Elements:

A) Based on properties, elements are classified into three. These are non-metals, metals and metalloids.

Metals:

- are electro positive elements
- are malleable and ductile
- are lustre and sonorous
- have high melting and boiling points
- are conductors
- are solids at room temperature except mercury which is liquid
- are found to the left bottom and middle part of the periodic table

Example: Aluminium, copper, gold, siver etc.

| Most of the elements are metals |

Non - metals

- are electron negative elements
- are not malleable and ductile
- have relatively low melting and boiling points
- occur in the three states of matter at room temperature.
- are non - conductors except graphite.
- are found to the right side of periodic table .

Example: carbon, oxygen, hydrogen etc.

Metalloids:

- are semi metals and semi non - metals
- are solids at room temperature
- are found at the boarder between metals and non-metals along the zigzag line.

Example: Boron, Germanium, Silicon, Antimony etc

B) Based on electron arrangements, elements are classified into three

1. Representative Elements

- They are 'A' group elements.
- They are s - block and p - block elements
- They compose metals, non-metals and metalloids.

2. Transition Elements

- They are 'B' group elements
- They are d - block elements
- All are metals
- They are found between IIA and IIIA
- They form colored and complex compounds
- They have variable valence number

Example

- Iron (II), Iron (III)
- Copper (I), Copper (II)

3. Inner Transition Elements

- They are also called **rare earth metals.**
- They are f - block elements
- They are all metals.
- They are found below the periodic table into two separate rows.
- The first row is called **the lanthanide series** and the second row is known as **the actinide series.**
- They are found in 6^{th} and 7^{th} periods.

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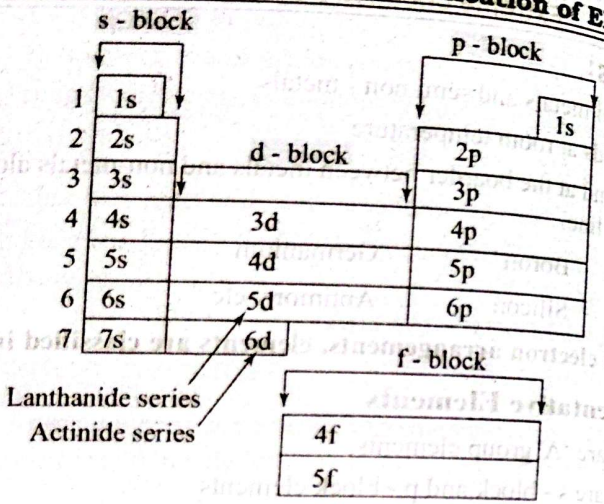


Fig 4.6 Classification of Element in Modern Periodic Table Based on Sub Shells

Illustrative Examples

23. Elements whose outer most shell electrons fall in the f-orbitals are called

A) Transition elements C) Metallic elements

B) Rare earth metals D) Alkaline earth metals

Solution

➥ Rare earth metals also called inner transition or f-block elements.

Answer: B

24. Given element X that has an electron configuration of $5s^2\ 4d^{10}\ 5p^1$. Which of the following is true regarding this element?

A) X is metallic element C) X is electro - positive element

B) X is d-block element D) X is a p-block element

Solution

Since the last electron entered in the p-subshell. It is a p-block element.

Answer: D

25. In which of the following blocks of the periodic table does an element with atomic number 14 located?

A) s - block B) d - block C) p - block D) f - block

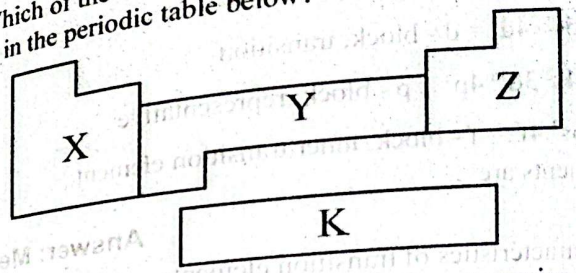
Solution

Write the electron configuration using sub-shells: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

Thus, it is p-block element.

Answer: C

26. Which of the following is true about positions indicated by X,Y,Z and K in the periodic table below?



A) All elements in which the last electron occupies the p-orbitals are placed in position Z.

B) 'B' group elements are placed in position Y and Z

C) All 'A' group elements are placed in position X

D) Elements in which the last electron goes to d - orbitals occupy position K.

Solution

➥ X = s - block elements ➥ Z = p - block elements

➥ Y = d - block elements ➥ K = f - block elements

Answer: A

27. Classify the following element as metals, non-metals and metalloid.

a) Aluminum c) Carbon e) Chlorine

b) Silicon d) Copper f) Boron

Solution

➥ a and d are metals

➥ b and f are metalloids.

➥ c and e are non-metals.

28. Classify the following elements as representative, transition or inner

transition

a) Sr (Z = 38)
c) Se (Z = 34)
b) Mo (Z = 42)
d) Ce (Z = 58)

Solution

Write electron configuration using sub shells

a) [Kr] $5s^2$ = s - block, representative
b) [Kr] $5s^1$ $4d^5$ = d - block, transition
c) [Ar] $4s^2$ $3d^{10}$ $4p^4$ = p - block, representative
d) [Xe] $6s^2$ $4f^2$ = f - block, inner transition element.

29. Most of the elements are _____.

Answer: Metal

30. List the main characteristics of transition elements.

Solution

- All are metals
- They have variable valencies.
- They form colored and complex compounds

4.4 General Properties With in Groups and Periods

1. Number of Valence Electrons

- They are electrons found in the valence shell of an atom.
- Number of valence electrons remain same down a group and increases across a period .
- Consider period 2 and group 4 elements below.

Group IA

$_1H$ 1
$_3Li$ 2, 1
$_{11}Na$ 2, 8, 1
$_{19}K$ 2, 8, 8, 1
$_{37}Rb$ 2, 8, 18, 8, 1
$_{55}Cs$ 2, 8, 18, 18, 8, 1
$_{87}Fr$ 2, 8, 18, 32, 18, 8, 1

All have same valence electrons

Period 2 elements

$_3Li$	$_4Be$	$_5B$	$_6C$	$_7N$	$_8O$	$_9F$	$_{10}Ne$
2, 1	2, 2	2, 3	2, 4	2, 5	2, 6	2, 7	2, 8

Valence electrons increase across a period

2. Number of Shells

- Number of shells increases down a group and remains same across a period from left to right

3. Nuclear Charge

- It is the total positive charge of the atom.

Example: $_{12}Mg$

Nuclear charge = + 12

Nuclear charge increases down a group and across a period.

4. Effective Nuclear Charge (Z_{eff})

- It is the difference between atomic number and inner electrons

$$Z_{eff} = Z - S$$

Z - atomic number

S - Number of inner electrons

Example: Determine Z_{eff} of silicon (Z = 14)

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Write electron configuration and determine the number of the electrons (electrons out of the valence electrons)

2, 8, 4 or $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

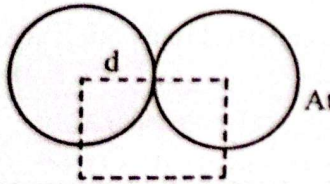
$S = 2 + 8 = 10$ or $2 + 2 + 6 = 10$

$Z_{eff} = Z - S = 14 - 10 = +4$

Z_{eff} remain same down a group and increases across period

5. Atomic size (Atomic radius)

- It may be defined as half the distance between two chemically bonded atoms in molecule.



Atomic radius (r) = $\frac{d}{2}$

For example, the distance between the nuclei in a Br_2 is 2.28A°.

Thus, atomic radius of Br = $\frac{d}{2} = \frac{2.28A^0}{2} = 1.14A^0$

- The atomic radius of an atom is very short. As a resurt, it is measured with angstrom (A°).

$1A^0 = 10^{-10}m = 10^{-8}cm$

- Atomic size **increases** down a group because number of shells increases.

- **Atomic size decreases across a period because of an increase in effective nuclear charge.**

- Atomic size **decreases** across a period because of an increase in effective nuclear charge.

Table 4.7 Periodic Variation of Atomic Radii of the Elements in the First three Periods in Nanometre (nm).

H 0.037							He 0.06
Li 0.134	Be 0.125	B 0.090	C 0.075	N 0.075	O 0.073	F 0.071	Ne 0.160
Na 0.154	Mg 0.145	Al 0.130	Si 0.118	P 0.110	S 0.102	Cl 0.099	Ar 0.190

The reduced attraction by the protons for the outer most electrons, caused by the core electrons is known as the shielding or screening effect.

Note:

1. The atomic size of positive ion is less than that of its neutral form
   Example: $Na^+ < Na$
2. The atomic size of negative ion is greater than its neutral form.
   Example: $F^- > F$
3. The more positive the cation in the same period the **smaller** is the size.
   Example: $Na^+ > Mg^{2+} > Al^{3+} > Si^{4+}$
4. The more negative the ion in the same period, the **larger** is the size.
   Example: $C^{4-} > N^{3-} > O^{2-} > F^-$

Illustrative Examples

31. Which of the following has the largest atomic radii?

A) $_{11}Na^+$ B) $_{12}Mg^{2+}$ C) $_{13}Al^{3+}$ D) $_{14}Si^{4+}$

Solution

All are found in the same period. The less positive the cation, the larger is the size.

Answer: A

32. Which of the following is the correct increasing order of the atomic size of the elements, Mg, B, K, Cl and C ?
(Z: Mg = 12, B = 5, C = 6, K = 19, Cl = 17)

B, C, Mg, Cl, K

A) B < C < Cl < K < Mg
B) C < B < Cl < Mg < K
C) Cl < C < B < K < Mg
D) Mg < K < Cl < C < B

Solution

- B and C are found in the same period, size C < B.
- Mg and Cl are found in the same period, size of Cl < Mg
- K is found in period 4 and has the largest atomic radius.
- C < B < Cl < Mg < K

Answer: B

33. The reason for the decrease in atomic radius as you go from left to right in a row of the periodic table is due to the:

A) Decrease in the effective nuclear charge
B) Increase in the number of neutrons
C) Increase in the nuclear charge
D) Addition of extra electron shells

Solution

An increase in nuclear charge across a period increases effective nuclear charge. This reduces the atomic size across a period.

Answer: C

34. Which of the following periodic properties increases down a group and across a period?

A) Nuclear charge
B) Atomic size
C) Valance electrons
D) Number of shells

Answer: A

35. Identify the shielding shells, shielding electrons, nuclear charge and effective nuclear charge for:

a) Al (Z = 13)
b) N (Z = 7)
c) Rb (Z = 37)

Solution

Write electron configuration

a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
- Shielding shells = K and L shells
- Number of shielding electrons = 2 + 2 + 6 = 10
- Nuclear charge = +13
- $Z_{eff} = Z\text{-}S = 13 - 10 = +3$

b) $1s^2\ 2s^2\ 2p^3$
- Shielding shell = K - shell
- Shielding electrons = 2
- Nuclear charge = +7
- $Z_{eff} = Z - S = 7 - 2 = +5$

c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^1$
- Shielding shell = K,L,M and N shell
- Number of shielding electron = 1
- $Z_{eff} = Z - s = 37 - 36 = +1$

36. Compare the atomic size of Al, Al^{1+}, Al^{2+}, Al^{3+}

Solution

$Al > Al^{1+} > Al^{2+} > Al^{3+}$

37. Determine the number of shells of an element with atomic number of 18.

Solution

Electron configuration = 2, 8,8

Number of shells = 3

38. Why atomic radius increases down a group?

Explanation: Because attraction force between the outer most electron and the nucleus decreases down a group

6. Ionization Energy (IE)

- It is the amount of energy required to remove an electron from isolated gaseous atom

$M(g) + IE \rightarrow M^{+}(g) + e^{-}$

- First ionization energy (IE_1) removes the first valence electron
- Second ionization energy (IE_2) removes the second valence electron
- Third ionization energy (IE_3) removes the third valence electron
- Fourth ionization energy (IE_4) removes the fourth electron

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$M(g) + IE_1 \rightarrow M^+(g) + e^-$

$M^+(g) + IE_2 \rightarrow M^{2+}(g) + e^-$

$M^{2+}(g) + IE_3 \rightarrow M^{3+}(g) + e^-$

$M^{3+}(g) + IE_4 \rightarrow M^{4+}(g) + e^-$

Note: $IE_1 < IE_2 < IE_3 < IE_4$

Factors Affecting Ionization Energy

a) Atomic Size

- As size increases ionization energy decreases and vice versal.

b) Effective Nuclear Change

- The smaller the effective nuclear charge, the lower the ionization energy.

c) Types of Electrons

- In the same energy level, the order of difficulty to remove an electron is: s - electron > p - electron > d - electron > f - electron

Example: 3s > 3p > 3d

d) Shielding /screening/ of Inner Electrons

- As the number of inner electrons increases, ionization energy decrease.

Example: IE of Li is greater than IE of Na because Na has more number of inner electrons.

- Fulfilled and half-filled orbitals are more stable and require higher IE to remove electron from them.

IE decreases down a group but increases across a period

- Metals are characterized by low IE but non-meals with high IE

Table 4.8 First Ionization energies of the first 18 elements in kilo joule per mole (KJ/mol)

1 H 1312							2 He 2372
3 Li 520	4 Be 900	5 B 801	6 C 1086	7 N 1402	8 O 1314	9 F 1618	10 Ne 2018
11 Na 496	12 Mg 738	13 Al 578	14 Si 789	15 P 1012	16 S 1000	17 Cl 1251	18 Ar 1521

Note: Exception across period 2 and 3

- IE of Be is greater than B due to fullfilled orbital in Be.
- IE of N is greater than O due to half - filled orbital in N.
- In period 3, IE of Mg > Al and P > S with similar reason as above.

7. Electro Negativity

- It is the power of an atom to attract shared electrons toward itself in a molecule or covalent compound.
- **Pauling** device a scale to compare the electro negativity of elements
- The element with the highest electro negativity value **(4)** is **fluorine**
- The element with the least electro negativity value is **cesium (0.7)**

Electro negativity decreases down a group and increases across aperiod

Table 4.9 Electro Negativities of First 18 Elements

1 H 2.1							2 He 0
3 Li 1.0	4 Be 1.5	5 B 2.0	6 C 2.5	7 N 3.0	8 O 3.5	9 F 4.0	10 Ne 0
11 Na 0.9	12 Mg 1.2	13 Al 1.5	14 Si 1.8	15 P 2.1	16 S 2.5	17 Cl 3.0	18 Ar 0

8. Electron Affinity (EA)

- It is defined as the energy lost when one mole of the gaseous atoms add one mole of electrons to from one mole of charged anions.

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$A(g) + e^- \rightarrow A(g)$, $EA < 0$

Example

$Cl(g) + \bar{e} \rightarrow Cl^-(g)$, $EA = -348 kJ/mol$

Electron affinity increases across a period and decreases down a group

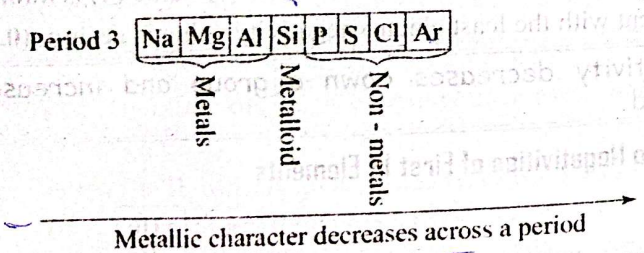
9. Metallic Character /Electro positivity/

- Metals have a tendency to lose electrons and form cations.
- The easier the metal can lose electrons, the more metallic it is.

Metallic character increases down a group and decreases across a period

C	Non - metal
Si	Metalloid
Ge	Metalloid
Sn	Metal
Pb	Metal

Metallic character increase down a group



Illustrative Examples

39. Which of the following elements has the lowest ionization energy? (Z: K = 19, Na = 11, Rb = 37, Cs = 55)

A) K B) Na C) Cs D) Rb

Solution

Down a group, ionization energy decreases

Answer: C

40. Increasing order of ionization energy for the elements Be, B, N and O is:

A) B < Be < O < N C) O < N < B < Be

B) Be < B < N < O D) N < O < Be < B

Solution

IE increases across a period, except in period 2 (Be > B, N > O) and period 3(Mg > Al, P > S)

➥ IE: B < Be < O < N

Answer: A

41. Which of the following is the correct order of electro negativities of elements C,N,O and F?

A) C > N > O > F C) F > O > N > C

B) C > F > O > N D) O > N > F > C

Solution

Electro negativity increases across a period.

➥ F > O > N > C

Answer: C

42. All of the following periodic properties decrease down a group. EXCEPT

A) Ionization energy C) Electro negativity

B) Metallic character D) Electron affinity

Answer: B

43. One of the following element has the highest electron affinity. Which one?

A) F B) Cl C) O D) N

Solution

Chlorine exceptionally has the highest negative electron affinity.

Answer: B

44. The element with the lowest electro negativity value is _____.

Answer: Cesium (Cs)

45. The elements that are characterized by low ionization energy and electro negativity are _____

Answer: Metals.

46. Refer the figure and answer the questions below it.

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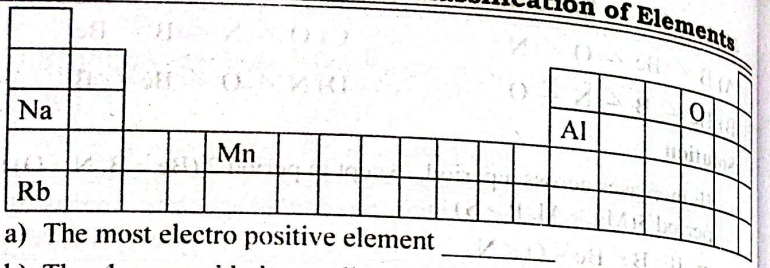

a) The most electro positive element _____

b) The element with the smallest size _____

c) Two elements that are found in the same

i. Group _____ and _____

ii. Period _____ and _____

d) The most electro negative element _____

e) The d - block element _____

Solution

a) Rb

b) O

c) i. Na and Rb
ii. Na and Al

d) O

e) Mn

47. Arrange the following in increasing order of ionization energy.

Na^+, Al^{3+}, Si^{4+}, Mg^{2+}

Answer: $Na^+ < Mg^{2+} < Al^{3+} < Si^{4+}$

4.5 The Diagonal Relationship and Advantages of Periodic Table

The diagonal relationship is shown by second and third period elements

Li → Mg

Be → Al

B → Si

Advantages of Periodic Table Include the Following

- It puts elements with similar properties in groups.
- It is useful in predicting properties of elements and their compounds
- It is used to predict formula of compounds
- It is used to determine position of elements
- It is used to refer name, symbol, atomic number and atomic mass of elements.

Illustrative Examples

48. A chemist can predict all of the following from the modern periodic table. **EXCEPT**

A) Formula of the compound that the elements can form

B) Chemical properties of elements

C) Colour of compounds

D) Position of elements

Answer: C

49. Which of the following pairs of elements have diagonal relationship?

A) Al and Mg

B) B and Si

C) Na and K

D) S and P

Answer: B

50. Classify oxides of period 3 elements as acidic, basic or amphoteric oxides

Solution

Na_2O and MgO ----- basic oxides

Al_2O_3 - amphoteric oxide

SiO_2, P_2O_5 and SO_2 - acidic oxides.

Solved Examples

51. What was Moseley's contribution to modern form of periodic table?

Solution

Henry Mosley discovered atomic number using X - ray which is unique characteristic of each element. He arranged elements based on atomic number. As a result, it becomes very easy to study the properties of elements.

52. What is the similarity between law of triad, law octaves and periodic law?

Solution

In the three laws, elements are arranged in increasing order of atomic masses.

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53. Who plotted the atomic volume Vs atomic mass graph to put elements having similar properties in the same position of the curve?

Solution

Lother Meyer

54. Explain the Mendeleev's short periodic table.

Explanation

- It consists 63 elements
- The elements are arranged in seven periods and eight groups.
- It contains vacant sites for undiscovered elements.

55. What is the cause of similarity properties in the same group?

Solution

Having same valence electrons

56. Write the valence shell configuration of

a) s - block elements
b) p - block elements
c) d - block elements
d) f - block elements

Solution

a) ns^{1-2}
b) $ns^2\ np^{1-6}$
c) $ns^{1-2}\ (n-1)d^{1-10}$
d) $ns^2\ (n-2)\ f^{1-14}$

57. Identify the four periods in which transition elements are found and the sub - shell that are filled with electrons.

Solution

Period 4 : 3d
Period 5 : 4d
Period 6 : 5d
Period 7 : 6d

58. Explain why there are no p - block elements in the first period of the periodic table?

Explanation

- The p -sub shell dose not exist in the first energy level.

59. Write the electron configuration and determine the atomic number of an element that corresponds to the following descriptions.

a) The noble gas in the fourth period
b) The group IA element in the fourth period

c) The group VIIA element in the fifth period

Solution

a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$
$Z = 2 + 2 + 6 + 2 + 6 + 2 + 10 + 6 = 36$

b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
$Z = 2 + 2 + 6 + 2 + 6 + 1 = 19$

c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^5$
$Z = 2 + 2 + 6 + 2 + 6 + 2 + 10 + 6 + 2 + 10 + 5 = 53$

60. Identify the following species with the descriptions

a) Smallest atomic radius in group IIIA
b) Most metallic element in group VA
c) Highest first ionization energy in IIA
d) Largest atomic size in period 3

Solution

a) Boron
b) Bismith
c) Berylium
d) Sodium

61. Arrange the following species in decreasing order of atomic size.

a) Be, C, Li, N
b) O^{2-}, Na^+, N^{3-}, F^-
c) B, Al, N, P
d) Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}

Solution

a) Li > Be > C > N
b) $N^{3-} > O^{2-} > F^- > Na^+$
c) Al > P > B > N
d) $Sr^{2+} > Ca^{2+} > Mg^{2+} > Be^{2+}$

62. Arrange the following in increasing order of ionization energy.

a) Mg, Mg^+, Mg^{2+}
b) Na, K, Rb, Li

Solution

a) $Mg < Mg^+ < Mg^{2+}$
b) Rb < K < Na < Li

63. Why noble gases have extremely low (almost zero) electron affinity?

Solution

This is because noble gases have stable fulfilled orbitals and low tendeney to accept electron.

64. Halogens have the highest electron affinity (highest negative value). Why?

Solution

Halogens need only one electron to be stable as a result, they have high tendency to gain an electron and form stable negative ions.

65. Classify the group IVA elements as non-metal, metalloids and metals.

Solution

- Carbon = Non - metal
- Silicon and germenium = metalloids
- Tin and lead = metals

66. Below is an incomplete periodic table of the first eighteen elements.

H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar

Which elements

a) are mono atomic and which diatomic?
Diatomic moocules: H_2, N_2, F_2, Cl_2

b) are alkali metals?

c) are unreactive towards other chemicals?

d) found in group VIA?

Solution

a) Mono atomic: He, Ne and Ar diatomic F and Cl

b) Li and Na

c) He, Ne and Ar

d) O and S

67. Explain why the ratio of $\frac{Fe^{2+}}{Fe^{3+}}$ is greater than one

Explanation

Because the ionic radius of Fe^{2+} is greater than that of Fe^{3+}

68. Distinguish between electro negativity and electron affinity and give their units of measurement if any.

Solution

- Electro negativity is the power of an atom to attract shared electrons toward itself.
- Electron affinity is the energy released when an electron negtivity and electron is added to isotated gaseous atom.
- Electro negativity has no unit but electron affinity has a unity of kilojoule per more (kJ/ mol).

69. What is periodic property?

Solution

- A periodic property is a property that repeats itself in the same manner with increase in atomic number.

70. Compare fluorine and chlorine in terms of electro negativity as well as electron affinity.

Solution

- Electron affinity : Cl > F
- Electro negativity : F > Cl

Supplementary Exercise

I. True/False items

1. Newland arranged elements based on atomic number
2. Elements in the same group have the same effective nuclear charge.
3. Period 4 and 5 have the same number of elements
4. The element with a valence shell electron configuration $2s^2\ 2p^5$ is a metal.
5. The period 6 inner transition series fills the 4 f- orbitals and are called actinide series.
6. An element with electron configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$, belong to period 3
7. Element in the same period have cosecutive atomic number.
8. Elements in the same group have similar physical properties.
9. Periodic table is important in predicting properties of compounds.
10. Period 2 is the shortest period that contains 8 elements.

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II. Matching Items

A	B
11. Group IA	A) Noble gases
12. Group IIA	B) Chalcogens
13. Group VB	C) Alkali metals
14. Group VIB	D) Halogens
15. Group VIIIA	E) Pnicogens
16. Group VIIA	F) Alkaline earth metals
	G) Boron family
	H) Carbon family

III. Multiple Choice Items

17. An element with atomic number 25 belongs to

A) s - block B) p - block C) d - block D) f - block

18. The property of element with atomic number 12 resembles that of an element with atomic number:

A) 11 B) 13 C) 1 D) 20

19. Which of the following elements would you expect to have chemical properties most similar to that of S?

A) P B) Cl C) Se D) Na

20. Which of the following are arranged correctly in order of increasing radius?

A) $F^- < Cl^- < Br^- < I^-$

B) $O^{2-} < F^- < Ne < Na^+$

C) $Ca^{2+} < Cl^- < K+ < Ar$

D) $Rb^+ < K^+ < Na^+ < Li^+$

21. Neglecting hydrogen atom, non-metals are located in:

A) s-block B) p-block C) d-block D) f-block

22. An element with a $4s^2\ 3d^2$ valence shell configuration is:

A) s-block metal

B) Found in period 2 of the periodie table

C) Found in group IVA of the periodic table

D) Definitely a transition metal

23. Which of the following elements has the greatest tendency to lose an electron if energy is supplied?

A) Potassium B) Calcium C) Oxygen D) Fluorine

24. Which of the following is Dobereiner's triad?

A) H, Li, Na B) Cl, Br, I C) Na, Mg, Al D) C, N, O

25. Which of the following elements has the smallest electron affinity?

A) Carbon B) Oxygen C) Neon D) Fluorine

26. In any group of the periodic table, elements.

A) become more reactive in moving down a group

B) tend to accept more readily as atomic number increases.

C) tend to lose electrons more readily as atomic size increases

D) become less reactive as atomic number increases

27. Which of the following set of elements have the same outer most electron configuration?

I. H, He, Be

II. H, Li, Be

III. H,Li, Na

IV. He, Ne, Ar

A. I and IV

B. III

C. I and III

D. IV

28. Which of the following is a noble gas?

A) Oxygen B) Ammonia C) Neon D) Hydrogen

29. Transition elements are:

I. Found between group IA and IIIA

II. All metals or semi - metals

III. All metals

IV. d-block element

A) III and IV B) I, III and IV C) I and III D) IV only

30. Which sequence of atomic numbers represent elements which have similar chemical properties?

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A) 4, 12, 20 B) 3, 4, 5 C) 11, 19, 20 D) 16, 17, 18

31. Which of the following groups consists of entirely metals?

A) IVA B) IA C) IIA D) VIA

32. Identify the most non-metallic element among the following.

A) Na B) Li C) Si D) P

33. Which of the following properties of elements decrease down a group of the period table?

A) Metallic property C) Ionization energy

B) Number of valence electrons D) Nuclear charge

34. Elements whose outer most shell electrons fall in the f-orbital are

A) Metallic elements C) Transition elements

B) Alkaline earth metals D) Rare earth metals

35. Among the following the particle with largest radius is:

A) Si^{4+} B) Al^{3+} C) Mg^{2+} D) Ca^{2+}

36. An element is located in the third period and group VA of the periodic table. Which of the following is true about this element?

A) It is a metallic element C) It has electron losing tendency

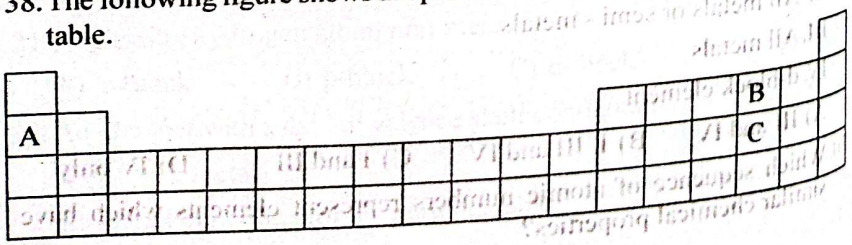
B) It forms a basic oxide D) It forms a trivalent ion

37. Which list of elements contains a metal, a metalloid, a non-metal and a noble gas?

A) Be, Si, Cl, Kr C) C, N, Ne, Ar

B) Na, Zn, Al, Sb D) K, Fe, B, F

38. The following figure shows the position of three elements in the periodic table.



Which of the trends in the properties of the elements is correct?

A) Ionization energy of B > C > A C) Atomic radius of C < A < B

B) Metallic character of A < C < B D) Electronegativity of A > B > C

39. Non - metals when compared to metals

A) Are less electrongative C) Have larger atomic size

B) Have greater ionization energy D) Are more electro positive

40. Elements not found in nature, synthesized in nuclear reactions and involving completion of the 5f electronic orbitals are :

A) the lanthanides C) the transition metals

B) the actinides D) the halogens

41. Give elements with their atomic radii (in A°) : X = 1.52, Y = 1.86, Z = 2.27 and W = 1.04. Which of the following is true about these elements?

A) Z has the highest Ionization energy

B) W is the least electro positive element

C) X has the greatest electron affinity

D) Y is the most electro negative element

42. Hydrogen can be placed in group IA because:

I. Can lose one electron in its reactions

II. Has one electron in its outer shell

III. Can gain one electron to complete its outer shell

IV. Has an oxidation number of + 2

A) IV B) IV and I C) II D) II and III

43. Lother Meyer plotted a graph showing variation of

A) Atomic volume with increase in atomic weight

B) Atomic radii with increase in atomic weight

C) Atomic volume with increase in atomic number

D) Atomic radii with increase in atomic number

44. The number of elements in the longest period is

A) 36 B) 8 C) 18 D) 32

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45. Among the elements A, B, C and D having atomic numbers 7, 8, 9 and 12 respectively, the element with smallest size and highest ionization energy is

A) A B) B C) C D) D

IV. Short answer items.

46. The most electronegative element is _____

47. Beryllium has diagonal relationship with _____

48. Determine which member has higher ionization energy.

a) Ca or Sr
b) N or F
c) Mg^{+} or Ca
d) Cr^{2+} or Cr^{3+}

49. Find the group and period of

a) Ar b) K

50. List the factors that affect ionization energy

Answer to Supplementary Exercise

Part I.

1. False	3. True	5. False	7. True	9. True
2. True	4. False	6. True	8. False	10. False

Part II.

11. C	12. F	13. E	14. B	15. A	16. D

Part III.

17. C	21. B	25. C	29. A	33. C	37. A	41. B	45. C
18. D	22. D	26. C	30. A	34. D	38. A	42. C	
19. C	23. A	27. B	31. C	35. D	39. B	43. A	
20. A	24. B	28. C	32. D	36. D	40. B	44. D	

Part IV

46. Fluorine

47. Aluminum

48.
a) Ca b) F c) Mg^{+} d) Cr^{3+}

49.
a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
Group = 2 + 6 = VIIIA
Period = 3

b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
Group = IA
Period = 4

50.
- Atomic size
- Effective nuclear charge
- Stability of orbitals
- Types of electrons

UNIT 5

Chemical Bonding

5.1 Introduction

Most atoms are unstable in their free form. They combine with each other or other type of atoms to form molecules or compounds. The force that holds atoms together in a molecule or compound is known as **chemical bonding** The bonds formed between atoms depend on the electron configuration of those atoms.

When atoms interact to form a chemical bond, only their outer regions come into contact. For this reason,when we study chemical bonding, we are usually concerned primarily with **valence electrons** (electrons found in the outer most shell of atoms)

Chemical bonding is attraction force that holds ions or atoms together in a compound or molecule. There are three types of chemical bonding, These are:

1. Ionic bonding
2. Covalent bonding
3. Metallic bonding

Lewis dot symbol - consists of the symbol of an element and one dot for each valence electron in an atom of the element.

Table 5.1 Lewis Dot Symbol for Periods 2 and 3 Elements.

Period 2	$Li^{\bullet}$	$\cdot Be\cdot$	$\cdot\dot{B}\cdot$	$\cdot\dot{C}\cdot$	$\cdot\ddot{N}\cdot$	$\cdot\ddot{O}\cdot$	$\cdot\ddot{F}\cdot$	$\cdot\ddot{Ne}\cdot$
Period 3	$Na^{\bullet}$	$\cdot Mg\cdot$	$\cdot\dot{Al}^{\bullet}$	$\cdot\dot{Si}\cdot$	$\cdot\ddot{P}\cdot$	$\cdot\ddot{S}\cdot$	$\cdot\ddot{Cl}\cdot$	$\cdot\ddot{Ar}\cdot$

Octet rule : Atoms lose, gain or share their valance electrons to have eight electrons in their outer most shell. Such atoms obey the octet rule.

All noble gases except **helium** obey the octet rule, without losing, gaining or sharing their valence electrons.

Example

$_{10}Ne : 2, 8$

$_{18}Ar : 2, 8, 8$

$_{36}Kr : 2, 8, 18, 8$

Metals lose their valence electrons to obey the octet rule.

Example

$Na \rightarrow Na^{+} + e^{-}$

$_{11}Na^{+} \rightarrow 2,8$

$Ca \rightarrow Ca^{2+} + 2e^{-}$

$_{20}Ca : 2, 8, 8$

Non-metals gain or share electrons to obey octet rule.

Example

$Cl + e^{-} \rightarrow Cl^{-}$

$2,8,7 \qquad 2,8,8$

Illustrative Examples

1. Why do atoms combine with chemical bonds to form compounds ? Because

A. To attain the lowest energy
B. To attain the highest energy
C. To be less stable
D. To be together

Solution

Most atoms are less stable when they are free. They combine together to lower their energy and become more stable

2. Which type of electrons participate in bond formation ?

A. For metals, all valence electrons
B. For non-metals, unpaired valence electrons
C. For non-metals, all valence electrons.
D. A and B are correct.

Answer: A

Solution

For metals the total valence electrons are lost to form cations while for non-metals unpaired valence electrons participate in bond formation.

Answer: D

3. How can atoms satisfy the octet rule?

Solution

Atoms satisfy the octet rule by:

- Losing valence electrons to other atoms
- Gaining valence electrons from other atoms
- Sharing their valence electrons with other atoms.

4. What type of force is a chemical bond?

Solution

It is an attraction force

5. Most substances are unstable in their elemental form. Why?

Explanation

Because, in their elemental form they have unfulfilled valence shell.

6. What type of electrons are involved in chemical bond formation?

Solution

Valence electrons

7. How many types of chemical bonding are there?

Solution

There are three types of chemical bonding

- Ionic bonding
- Covalent bonding

- Metallic bonding

8. During chemical bond formation, What type of change occurs?

Solution

A chemical change take place because of change in composition of substances.

5.2 Ionic Bonding

☑ **Ionic bonding** is an attraction force between positive and negative ions.
☑ **Positive ions or cations** are formed when atoms lose electron /s/.

Example

$$Ca \rightarrow Ca^{2+} + 2e^-$$

☑ **Negative ions or anions** are formed when atoms gain electron /s/.

Example

$$S + 2e^- \rightarrow S^{2-}$$

Ionization is a process of forming ions.

☑ Ionic bonding is also known as **electrovalent bond** because it is an attraction force between cations and anions.

Formation of Ionic Bonding

☑ **Ionic bonding** is formed by complete transfer of electron /s/ from metal to non - metal.
☑ It is formed between elements that have low ionization energy (metals) and high electron affinity (non-metals).

Example

- Formation ionic bond between

a. Calcium and oxygen

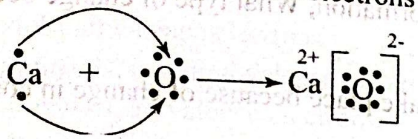
- First, write electron configuration and determine valence electrons

$_{20}Ca$: 2, 8, 8, 2, valence electrons = 2

$_{8}O$: 2, 6, valence electrons = 6

- Then, draw a Lewis dot symbol for each element and show

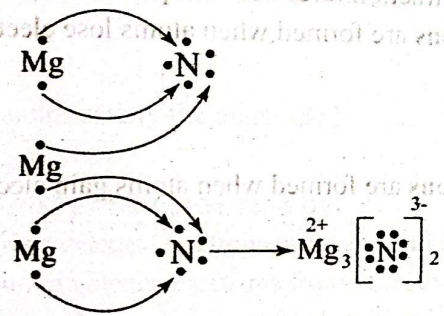
electron transfer from metal to non-metal. The charge of the metal and non-metal is equal to the electrons lost and gained respectively.



b. Magnesium and nitrogen

$_{12}Mg$: 2, 8, 2, valence electrons = 2

$_{7}N$: 2, 5, valence electrons = 5



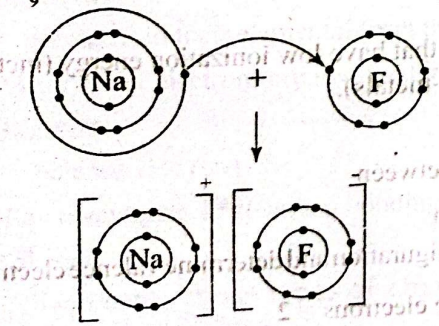
Formation of Ionic Bonding Using Bohr Diagram

Example

Between sodium and fluorine

$_{11}Na$: 2, 8, 1

$_{9}F$: 2, 8, 7



Lewis Structure of Ionic Compounds

To draw the Lewis structure of ionic compounds,

(i) Write the charge of cation as superscript

(ii) Write the valence electrons of the anion using closed interval and place its charge as superscript.

Example

Draw the Lewis structure of

a. Sodium chloride (Na Cl)

b. Aluminumoxide (Al_2O_3)

Solution

a) NaCl = $Na^{+}[\dot{\underset{\cdot\cdot}{Cl}}\cdot]^{-}$ b) Al_2O_3 = $Al_2^{3+}[O]_3^{2-}$

Illustrative Examples

9. The attraction force between cation and anion is

A. Ionic bond
B. electrovalent bond
C. Covalent bond
D. A and B

Answer: D

10. The process in which a neutral atom loses or gains its valence electrons to form ion is known as _____ .

A. Ionic bonding
B. Ionic bond formation
C. Ionization
D. Lewis structure.

Answer: C

11. Which of the following is not true about a cation?

A. It has more protons than electrons
B. Its size is smaller than its neutral form.
C. It is formed when its atom form loses electrons.
D. It has negative charge.

Solution

Cation - is positive ion

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- p > e
- Has smaller size
- Formed when a neutral atom loses electrons.

Answer: D

12. In general, ionic bond is formed between _____ and _____.

Solution

Metals and non-metals.

13. Explain the formation of ionic bonding

Solution

Ionic bonding is formed by complete transfer of electrons from metals to non-metals.

14. Why metals lose and non-metals gain electrons to form ions?

Explanation

Non-metals are more electro - negative elements than metals. They have a better tendency to attract or gain electrons than metals. They have also more valence electrons than metals.

15. Draw the Lewis structure of the following ionic compounds.

a. K_2O b. MgS c. Ca_3P_2

Solution

a. $K_2^{+}[:\ddot{O}:]^{2-}$

b. $Mg^{2+}[:\ddot{S}:]^{2-}$

c. $Ca_3^{2+}[:\ddot{P}:]_2^{3-}$

16. What type of bond is formed between elements having electron configuration of $[He]2s^1$ and $[He]\ 2s^2\ 2p^5$ respectively?

Solution

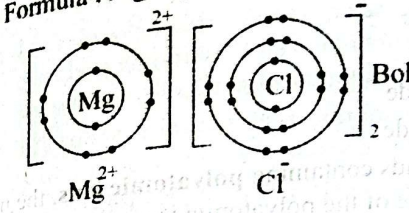
Metals usually have 1 to 3 valence electrons and non-metals from 4 to 8 valence electrons. Thus, the first element is metal and the second is non-metal. The type of bond formed between metal and non-metal is ionic bond.

17. Draw Bohr's diagramatic representation for

a. Magnesium chloride b. Sodium oxide

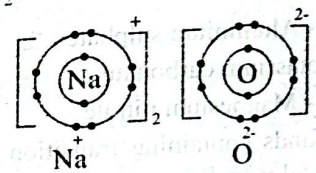
Solution

a. Formula : $MgCl_2$



b. Sodiumoxide: Na_2O

Bohr's diagram:



18. Write the formula of the ionic compounds represented by the Lewis structures below.

a. $Na^{+}[:\ddot{Cl}:]^{-}$ c. $Mg_3^{2+}[:\ddot{N}:]_2^{3-}$

b. $Mg^{2+}[:\ddot{O}:]^{2-}$ d. $Al^{3+}[:\ddot{N}:]^{3-}$

Solution

Simply omit the charges and dots

a. NaCl b. MgO c. Mg_3N_2 d. AlN

Naming of Ionic Compounds

- ☑ Ionic compounds are the results of ionic bonding
- ☑ For naming binary ionic compounds, the name of the metals remain same but that of the non-metals ends with the **suffix - ide**

Oxygen - Oxide	Phosphorus - phosphide
Hydrogen - hydride	Sulphur - sulphide
Nitrogen - nitride	Halogen - halide

Example

- CaO - calcium oxide
- NaI - sodium iodide
- Ba_3N_2 - Bariumnitride
- NaH - sodium hydride

☑ For naming ionic compounds containing **polyatomic ions**, the name of metal followed by the name of the polyatomic ion.

Example

- $Al_2(SO_4)_3$ - Aluminium sulphate
- K_2CO_3 - Potassium carbonate
- $Mg(NO_3)_2$ - Magnesium nitrate

☑ For ionic compounds containing transition metals, state the valence number of the metal with Roman number in bracket, after the name of the metal.

Example

Cu_2O - copper (I) oxide
CuO - copper (II) oxide
$FeSO_4$ - Iron (II) sulphate
$Fe_2(SO_4)_3$ - Iron (III) sulphate

Properties of Ionic Compounds

Ionic compounds have the following properties.

1. **They are aggregates of positive and negative ions**

For example, consider Na Cl

Na^+ Cl^- Na^+ Cl^- Na^+ Cl^- Na^+ Cl^-
Cl^- Na^+ Cl^- Na^+ Cl^- Na^+ Cl^- Na^+
Na^+ Cl^- Na^+ Cl^- Na^+ Cl^- Na^+ Cl^-
Cl^- Na^+ Cl^- Na^+ Cl^- Na^+ Cl^- Na^+
Na^+ Cl^- Na^+ Cl^- Na^+ Cl^- Na^+ Cl^-

2. **They are usually soluble in polar solvent (water) but insoluble in non-polar solvent such as benzene, Carbon tetrachloride, acetone, ethanol etc.**
   - The solubility of ionic compounds follows the " **like dissolves like**" principle.
   - Like dissolves like means, polar dissolves polar and non-polar dissolves non-polar. Ionic compounds are polar that is why they are soluble in polar solvent but insoluble in non-polar solvents.
   - The solute -solvent interaction force responsible for the solubility of ionic compound in water is **ion - dipole force.**
3. **They have high melting and boiling points.**
   - This is due to the strong electro static force of attraction between oppositely charged ions.
   - As the size of the ions decreases, the melting and boiling points of ionic compounds increase because of an increase in bond strength.

Example

Melting point of: NaF > NaCl > NaBr > NaI:

Reason size $F^- < Cl^- < Br^- < I^-$

4. **The are hard and brittle**
   - Due to the strong attraction force between cations and anions, ionic compounds are hard.
5. **They have high density**
   - Density is mass per unit volume. The ions are very close to each other. Thus, there are more ions per unit volume. That is the reason for high density of ionic compound
6. **They are non-volatile compounds.**
   - **Volatility** is the tendency of a substance to vaporize at room temperature. Since ionic compounds are formed because of the strongest ionic bond, they are non-volatile.
7. **Ionic compounds are crystalline solids**
   - They have regular arrangement of particles.
   - They have fixed melting and boiling points.
8. **They are electrolytes.**
   - Ionic compounds transmit electricity in molten or aqueous state.

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They do not transmit electricity in solid state because the ions are not free to carry charge.

Example

$NaCl(1)$ - molten
$NaCl\,(aq)$ -aqueous solution } Conductor

$NaCl\,(s)$ - Solid, non - conductor

Illustrative Example

19. The formula of iron (11) sulphite is

A. FeS B. FeS_2 C. $FeSO_3$ D. $Fe_2(SO_3)_3$

Solution

$Iron(II) = Fe^{2+}$

$Sulphite = SO_3^{2-}$

$Formula = Fe^{2+}SO_3^{2-} = Fe_2(SO_3)_2 = FeSO_3$

when simplified by 2 **Answer: C**

20. All of the following are ionic compounds. EXCEPT

A. CaO B. HCl C. LiF D. Na_2O

Solution

Ionic compounds contain metal and non - metal or ammonium ion (NH_4^+) and non-metal. HCl is a compound of two non-metals, hydrogen and chlorine Thus, it is not an ionic compound.

Answer: B

21. Ionic compounds are generally

A. Soft molecular solids
B. rigid molecular solids
C. Volatile molecular liquids
D. Closely packed dense solids.

Solution

Ionic compounds are hard, brittle, non-volatile, crystalline and closely packed dense solids.

Answer: D

22. Which of the following does not transmit electricity?

A. KCl (s) B. KCl (aq) C. KCl (1) D. Cu (s)

Solution

Ionic compounds are non-conductors in solid state but metal are conductors in solid state.

Answer: A

23. Most ionic compounds are soluble in

A. C_6H_6 B. CCl_4 C. C_2H_5OH D. H_2O

Answer: D

24. Arrange the following ionic compounds in order of increasing melting point. LiF, KF, NaF, RbF

Solution

Melting point deccases with increasing size of ions (cation or anions)

➥ RbF ∠ KF ∠ NaF ∠ LiF

25. What are the charge carriers in ionic conduction?

Solution

Positive and negative ions (cations and anions)

26. Ionic compounds are crystalline solids. Write the difference between crystalline and amorphous solids.

Solution

Crystalline Solid	**Amorphous Solid**
➥ Regular pattern of particles	➥ Irregular pattern of particles
➥ Has fixed melting and boiling points	➥ Has variable melting and boiling points
Example	**Example**
Ionic compounds, metals	Glass, rubber, plastics.

27. Write the name of the following ionic compounds.

a. RbBr c. $ZnCl_2$ e. CrO
b. SrO d. Cr_2O_3 f. $Al_2(SO_4)_3$

Solution

a. Rubidium bromide b. Strontium oxide

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c. Zinc chloride
e. Chromium (11) oxide
d. Chromium (111) oxide
f. Aluminum sulphate

28. Write the formula of an ionic compound formed between

a. Mg and P
b. Li and N
c. Al and F
d. Ba and S

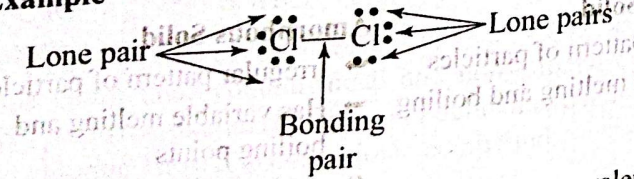
Solution

a. $Mg^{2} \times P^{3} = Mg_3P_2$
c. $Al^{3} \times F^{1} = AlF_3$
b. $Li^{1} \times N^{3} = Li_3N$
d. $Ba^{2} \times S^{2} = BaS$

5.3 Covalent Bonding

- ☑ It is also called **electron pair bond.**
- ☑ It is an attraction force between **shared electrons** and the nuclei of bonded atoms.
- ☑ It is formed by **sharing of electrons** between elements that have similar electro negativities.
- ☑ The shared electrons localized between the bonded atoms are called **bonding pairs (bp)**
- ☑ The unshared electrons that do not participate in bond formation are called **non-bonding pairs or lone pairs (Lp)**

Example



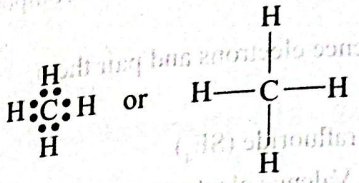
- ☑ **A Lewis structure** is a representation of a covalent bonding using Lewis dot symbol in which shared electron pairs are shown either as lines or as pairs of dots between the bonded atoms and lone pairs as pairs of dots are shown on individual atoms.

Based on the number of Electron Pairs shared, Covalent Bond is Divide in to Three.

(i) Two atoms held together by one electron pair are said to be joined by **a single covalent bond.**

Example

H_2 : H : H or H — H one single covalent bond

CH_4 :  4 single covalent bonds

Note: If bonding pairs are represented with dots, the formula is know as dot formula. If they are represented with lines, the formula is known as line formula.

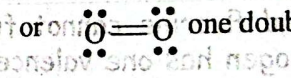
Example

For Cl_2 :

Dot formula Line formula

(ii) If two atoms share two pair of electrons, the covalent bond is called **double covalent bond.**

Example

O_2 :  one double covalent bond

CO_2 : two double covalent bonds

(iii) **A triple covalent bond** arises when two atoms share three pairs of electrons.

Example

N_2 : :N:::N: or :N≡N:

C_2H_2 : H—C≡C—H or H:C:::C:H

Note: i. Bond strength: Triple bond > double bond > single bond

ii. Bond strength: Single bond > double bond > Triple bond

Lewis Structure for Covalent Compounds

To draw the plausible Lewis structure of covalent compounds follow the steps below:

1. Determine the total valence electrons and pair them.

Example

Consider sulphur tetrafluoride (SF_4)

S : Z = 16, 2, 8, 6, Valence electrons = 6

F : Z = 9, 2, 7, Valence electrons = 7

$\Rightarrow$ total valence electrons of F = $4 \times 7 = 28$

Total valence electrons = Valence electron S+ Valence electrons of F

$= 1 \times 6 + 4 \times 7 = 34$

$$\text{Pairs} = \frac{34}{2} = 17$$

2. Identify the central atom and draw the skeletal structure.

Centeral atom is the one with larger size or more ectropositive

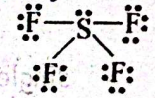
➥ For SF_6, S is centeral atom

Note: Hydrogen and fluorine cannot function as central atom because hydrogen has one valence electron smallest size and fluorine is the most electronegative atom

➥ Skeletal structure of SF_4 : F—S—F (F, F)

➥ There should be gaps between the bonding and lone pairs because of the electron - electron repulsion force.

3. Fulfill the octet of terminal (surrounding) atoms except hydrogen and put any remaining pair of electrons on the central atom.



bp = 4

Lp = 13

Thus, the above structure is the lewis structure of SF_4.

4. You may make double or triple bond to fulfill the octet of the centeral atom for some molecules.

Example

SO_3

$_{16}S : 2, 8, 6$

S : $1 \times 6 = 6$

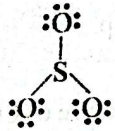
O : $3 \times 6 = 18$

$_8O$: 2, 6

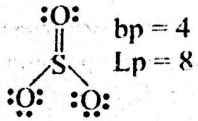
➥ Total valence electrons = 6 + 18 = 24

➥ Pairs = $\frac{24}{2} = 12$

➥ Central atom : S because it has larger size than O



No electron pair left and S is out of octet. Therefore, use a lone pair of electrons from either of the three oxygen atoms and make a double bond with sulphur



bp = 4

Lp = 8

Illustrative Examples

29. Which of the following type of covalent bond has the highest bond energy?

A. A triple bond
B. A double bond
C. A single
D. None

Solution

Bond energy and strength are directly related. That is, the stronger the bond, the higher is the bond energy. A triple bond is the strongest bond and has the highest bond energy.

Answer: A

30. In the Lewis structure of NF_3, nitrogen is surrounded by _____ bonding pairs and _____ lone pairs.

A. 3 , 1
B. 1 , 3
C. 2 , 2
D. 4 , 0

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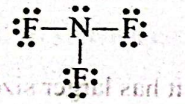
Solution

$_7N : 2, 5 : 1 \times 5 = 5$

$_9F : 2, 7 : 3 \times 7 = 21$

Total valence electrons = 5 + 21 = 26

$$\text{Pairs} = \frac{26}{2} = 13$$

Lewis structure : 

Thus , N is surrounded by 3 bp and 1Lp.

Answer:A

31. How many lone pairs are in the lewis structure of CCl_4?

A. 4 B. 10 C. 8 D. 12

Solution

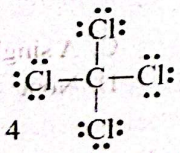
CCl4 : $_6C : 2, 4, \quad 1 \times 4 = 4$

$_{17}Cl : 2, 8, \quad 4 \times 7 = 28$

Total valence electrons = 4 + 28 = 32

$$\text{Pairs} = \frac{32}{2} = 16$$

centeral atom: C

Lewis structure : 

Lp = 12 , bp = 4

Answer: D

32. Determine the total valence electron of the following

a. CO_2 b. N_2H_4 c. H_3O^+ d. NO_3^-

Solution

a. CO_2 : $_6C : 2, \underline{4}, 1 \times 4 = 4$

$_8O : 2, \underline{6}, 2 \times 6 = 12$

Total valence electrons = 4 + 12 = 16

b. N_2H_4 : $_7N : 2, \underline{5}, 2 \times 5 = 10$

$_1H : 1, \quad 4 \times 1 = 4$

Total valence electrons = 10 + 4 = 14

c. H_3O^+ : $_1H : 1, \quad 3 \times 1 = 3$

$_8O : 2, \underline{6}, 1 \times 6 = 6$

charge = -1

Total valence electrons = 3 + 6 - 1 = 8

Note: For polyatomic ions, subtract the magnitude for cations and add the magnitude of the charge for anions.

d. NO_3^- : $_7N : 2, \underline{5}, 1 \times 5 = 5$

$_8O : 2, \underline{6}, 3 \times 6 = 18$

Charge = + 1

Total valence electrons = 5 + 18 + 1 = 24

33. Hydrogen and fluorine cannot serve as centeral atom. Explain why?

Explanation

Reason : Hydrogen has only one valence electron .

Fluorine is the most electronegative element so far known.

34. When you draw a Lewis structure of a covalent compound, you have to leave a gap between bonding pairs and lone pairs. why ?

Explanation

- This is due to the electron-electron repulion forces between bonding and lone pairs.
- The order of repulsion is:

Lp - Lp > Lp - bp > bp -bp.

35. Draw Lewis structure and determine the number of bonding and lone pairs for the following covalent molecules or molecular ions.

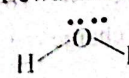
a. H_2O c. BH_3 e. NH_4^+

b. HCN d. SF_6 f. ClO_3^-

Solution

a. H_2O : $_1H : 1 = 2 \times 1 = 2$

$_8O : 2, 6 : 1 \times 6 = 6$

Lewis structure 

bp = 2

Lp = 2

Total valence electrons = 2 + 6 = 8

central atom : O

b. HCN:

$_1H : 1 , 1 \times 1 = 1$

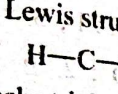
$_7N : 2 , 5 , 1 \times 5 = 5$

$_6C , 2 , 4 , 1 \times 4 = 4$

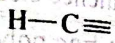
Total valence electrons = 1 + 5 + 4 = 10

central atom = C

Lewis structure



make triple bond to fulfill the octet of carbon



bp = 4 , Lp = 1

c. BH_3

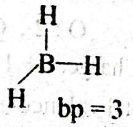
$_5B , 2 , 3 , 1 \times 3 = 3$

$_1H , 1 , 3 \times 1 = 3$

Total valence electrons = 3 + 3 = 6

centeral atom = B

Lewis structure



bp = 3

Lp = 0

d. SF_6

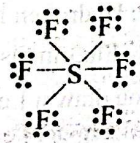
$_{16}S : 2 , 8 , \underline{6} : 1 \times 6 = 6$

$_9F : 2 , \underline{7} : 6 \times 7 = 42$

Total valence electrons = 6 + 42 = 48

centeral atom : S

Lewis structure



bp = 6

Lp = 18

e. NH_4^+

$_7N : 2 , \underline{5} : 1 \times 5 = 5$

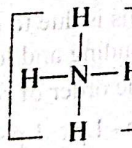
$_1H : 1 : 4 \times 1 = 4$

charge = -1

Total valence electrons = 5 + 4 - 1 = 8

centeral atom : N

Lewis structure



bp = 4

Lp = 0

Note: For molcular ions use closed interval bracket and put the change as superscript.

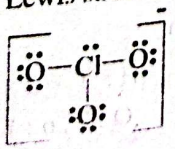
f. ClO_3^-

$_{17}Cl : 2 , 8 , \underline{7} : 1 \times 7 = 7$

$_8O : 2 , 6 : 3 \times 6 = 18$

charge = + 1

Lewis structure



bp = 3

Lp = 10

Total valence electrons = 7 + 18 + 1 = 26 bp = 3

centeral atom : Cl Lp = 10

36. Determine the number of lone pairs on the centeral atom of the following molecules or molecular ions.

a. H_3O^+ b. ICl_2^- c. BrF_3 d. $SiCl_4$

Solution

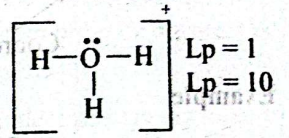
a. $H_3O^+ : H : 3 \times 1 = 3$

$O : 1 \times 6 = 6$

charge = $\frac{-1}{8}$

pairs = $\frac{8}{2} = 4$

Lewis structure



Lp = 1

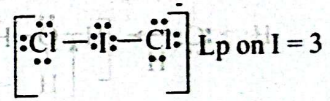
Lp = 10

b. $ICl_2^- : I : 1 \times 7 = 7$

$Cl : 2 \times 7 = 14$

chage = $\frac{+1}{22}$

pairs = 11

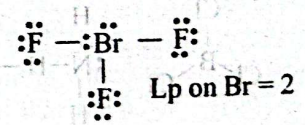


Lp on I = 3

c. $BrF_3 : Br : 1 \times 7 = 7$

$F : 3 \times 7 = \frac{21}{28}$

pairs = 14

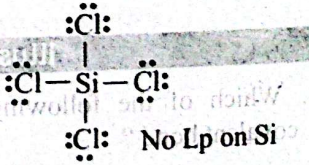


Lp on Br = 2

d. $SiCl_4 : Si = 1 \times 4 = 4$

$Cl : 4 \times 7 = \frac{28}{32}$

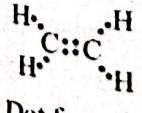
pairs = 16



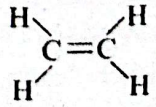
No Lp on Si

37. Draw the dot and line formula of ethane (C_2H_4) and list the type of covalent bonds.

Solution



Dot formula



Line formula

C_2H_4 has four single covalent bonds and one double covalent bond

Coordinate Covalent

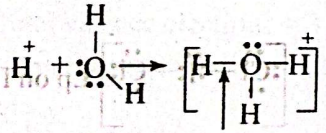
- It is also callad **dativ bond**
- It is a type of covalant bond in which the shared electrons are contributed by one of the bonded atoms.

$X + :Y \longrightarrow X:Y$ or $X-Y$

Coordinate covalent bond

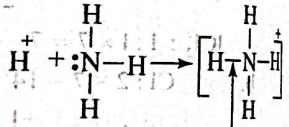
Example

a) $H^+ + H_2O \longrightarrow H_3O^+$



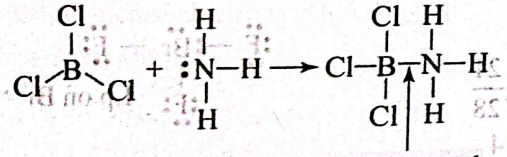
Dative bond

b) $H^+ + NH_3 \longrightarrow NH_4^+$



Dative bond

c) $BCl_3 + NH_3 \longrightarrow Cl_3BNH_3$



Dative bond

Illustrative Examples

38. Which of the following hypothetical reactions forms coordinate covalent bond ?

A. $G: + H^+ \longrightarrow [G:H]^+$

B. $E\cdot + F\cdot \longrightarrow [E^+:F^-]$

C. $C^+ + D^- \longrightarrow C^+D^-$

D. $A\cdot + \cdot B \longrightarrow A:B$

Solution

$G: + H^+ \longrightarrow [G:H]^+$ (donor acceptor)

$E\cdot + \cdot F \longrightarrow E^+F^-$ (ionic bond)

$A\cdot + \cdot B \longrightarrow E:F$ (Covalent bond)

Answer: A

39. Explain the difference between covalent and coordinate covalent bond.

Solution

In coordinate covalent bond, the shared electrons are obtained from one of the bonded atom but in covalent bond both atoms share electrons.

40. Write the other name of coordinate covalent bond.

Solution

Dative bond or donor - acceptor bond.

Polarity in covalent molecules

- ☑ **Polarity** is unsymmetrical charge distribution.
- ☑ **A polar covalent bond** is a covalent bond in which the shared pair of electrons is strongly attracted to one atom than the other.
- ☑ The atom with a higher electronegativity will have a partial negative charge [δ-] and the other a partial positive charge (δ^+).

Example

In HCl, Cl is more electronegative

$H^{\delta+}—Cl^{\delta-}$

- Polar bonds unless symmetrically arranged, produce polar molecules.
- A polar molecule is sometimes called a dipole.
- A **dipole** has asymmetrical charge distribution
- Partial charges within a molecule are indicated by delta (δ) .

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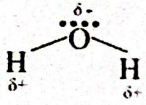
Example



- **Dipole moment** (μ) used to describe the strength of a dipole. It is the distance that separate them.

$$\mu = Q \cdot d$$

Q - is the size of partial charge in coulombs

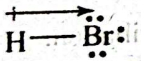
d - distance that separates the partial charges in meter

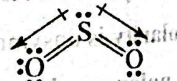
- The SI unit of μ is coulomb. meter (C . m)
- The other unit of μ is debye (D)

$$1D = 3.338 \times 10^{-30}\ C.\ m$$

- The higher the value of μ, the stronger is the bond.
- **Bond polarity** can be indicated by using an **arrow** ($\longrightarrow$) with short perpendicular vertical line at its tail. Its tail is placed on more positive atom and its head on more negative atom.

Example

a) HBr : 

b) SO_2 : 

Non-polar covalent bond: is formed when there is equal distribution of electrons between the bonded atoms.

- It is formed between elements that have almost similar electro-negativities.

Example

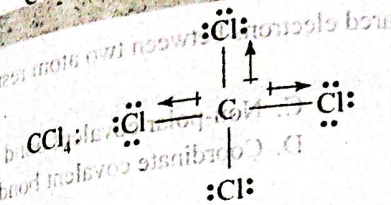
H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2, CH_4 etc.

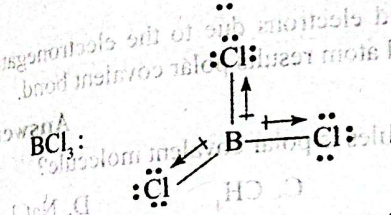
- Covalent molecules with zero dipole moment are called **non-polar covalent molecules.**

Example

H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2, CH_4, CO_2, CCl_4, BCl_3, etc.

CO_2 : $\mu = 0$ Because the sum of the two arrows is zero

CCl_4 :  $\mu = 0$

BCl_3 :  $\mu = 0$

Note: CO_2, CCl_4 and BCl_3 are **symmetric** compounds with polar bonds.

Table 5.2 Electronegativity Difference and Type of Bond for Most Compounds

Electronegativity Difference	Type of Bond
0 - 0.5 or [0 , 0.5)	Non - polar covalent bond
0.5 - 2 . 0 or [0.5 - 2.0)	Polar covalent bond
Greater that or equal to 2.0	Ionic bond

Example

a. CH_4 Electronegativity of (EN):

C = 2.56

H = 2.20

Electronegativity (EN.d) difference = 2.56 - 2.20 = 0.36

Thus, C - H bond is non-polar bond

b. HF : EN : H = 2.20, F = 4.00

EN. d = 4.00 - 2. 20 = 1.80

Thus, H - F bond is polar bond

c. NaCl : EN : Na = 0.96, Cl = 3.00

EN. d = 3.00 - 0.96 = 2.04

Thus, the bond between Na and Cl is ionic bond.

Illustrative Examples

41. Unequal distribution of the shared electrons between two atom results in the formation of:

A. Ionic bond
B. Polar covalent bond
C. Non-polar covalent bond
D. Coordinate covalent bond

Solution

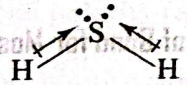
Unequal distribution of shared electrons due to the electronegativity difference between the bonded atom results polar covalent bond.

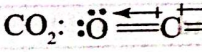
Answer: B

42. Which of the following molecules is polar covalent molecule?

A. H_2S B. CO_2 C. CH_4 D. NaCl

Solution

H_2S:  $\mu \neq 0$ It is polar covalent molecule

CO_2:  $\mu = 0$ It is non - polar covalent molecule

CH_4: has no bond polarity ($\mu = 0$) as a result it is non-polar

NaCl is an ionic compound

Answer: A

43. The most polar bond among the following is

A. Si ----- O B. Si ----- F C. Si ----- Cl D. Si ----- Br

Solution

The higher the electro negativity difference between the bonded atoms, The more polar is the bond

Answer: B

44. Which of the following contains both ionic and covalent bonds?

A. SF_4 B. NaCl C. KCN D. CF_4

Solution

Ionic and **covalent bonds** are found in compounds than contain metal or ammonium ion and polyatomic ions

$\left(OH^-, CN^-, NO_3^-, NO_2^-, SO_4^{2-}, CO_3^{2-}, PO_4^{3-}, HCO_3^-, HSO_4^-, H_2PO_4^-, HPO_4^{2-}, \text{etc}\right)$ = are polyatomic ions

Answer: C

45. Which of the following molecule is non-polar with polar bond?

A. H_2O B. NF_3 C. CH_4 D. CCl_4

Solution

Symmetric molecules with different electro negativity values are non-polar covalent molecules with polar bonds.

Example

CO_2, CCl_4, BCl_3, PCl_5, SF_6

Answer: D

46. Classify the following covalent compounds as polar or non-polar covalent compounds.

a. CS_2 d. C_6H_6 g. SiF_4
b. H_3O^+ e. NH_3 h. PCl_5
c. $CHCl_3$ f. C_7H_8 i. SF_2

Solution

- b, c, e and i are polar covalent compounds.
- a, d, f, g and h are non-polar covalent compounds.

47. Classify the following bonds as polar covalent, non-polar covalent or ionic based on the values of electro negativity.

a. NaF ; Na = 0.96, F = 4.00
b. NH_3 ; N = 2.81, H = 2.20
c. $CH_3 - CH_3$, C = 2.56 H = 2.20

Solution

a. Electro negativity difference = 4.00 - 0.96 = 3.04
   ∴ The bond in NaF is ionic
b. Electronegtivity difference = 2, 81 - 2, 20 = 0.61
   ∴ The N - H bond is polar covalent bond
c. Electronegativity difference (C - H) = 2. 56 - 2. 20 = 0.36

CamScanner

Electronegativity difference (C - C) = 2. 20 - 2. 20 = 0

∴ Both C - H and C - C bonds are non-polar covalent bonds.

48. Indicate the bond polarity of the following bonds

a. H - O b. N - H c. S - F d. Cl - I

Solution

a. $H \xrightarrow{+} O$ b. $N \xleftrightarrow{+} H$ c. $S \xrightarrow{+} F$ d. $Cl \xleftrightarrow{+} I$

49. List the types of bonds present in ammonium chloride (NH_4Cl)

Solution

- Between NH^+_4, and Cl^- = ionic bond
- In NH^+_4, between NH_3 and H^+ = coordinate covalent bond
- In NH_3, between N and H = polar covalent bond

50. Which bond is stronger? polar covalent or non - polar covalent bond.

Solution

Polar covalent bond.

Exception to octet rule in covalent compounds

- If there is no eight electrons around the centeral atom, such type of covalent compounds violate the octet rule. This happens in three cases.

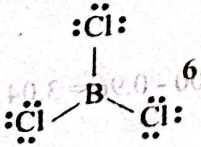
1. Molecules Below Octet

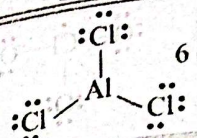
- Such molecules have below eight electrons around the central atom.

Example

Be Cl_2, BCl_3, Al Cl_3 etc

$BeCl_2$: :Cl—Be—Cl: 4 electrons on Be

BCl_3:  6 electrons on B

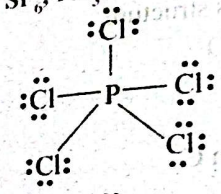
$AlCl_3$:  6 electrons on Al

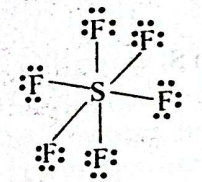
2. Molecules Above Octet

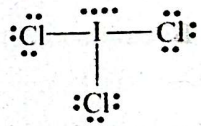
These molecules have above 8 electrons around the central atom.

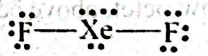
Example

PCl_5, SF_6, ICl_3, XeF_2, etc

PCl_5:  10 electrons on P

SF_6:  12 electrons on S

ICl_3:  10 electrons on I

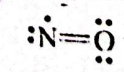
XeF_2:  10 electrons on Xe

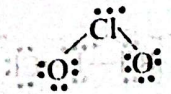
3. Molecules with Odd Number of Electrons

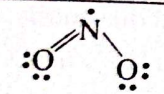
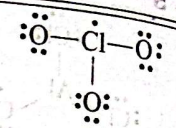
All covalent molecules with odd number of electrons violet the octet rule because eight is even number

Example

NO, NO_2, ClO_2, ClO_3, etc

NO: 

ClO_2: 

NO_2:  ClO_3: 

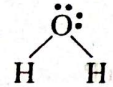
Illustrative Examples

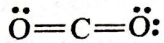
51. Which of the following molecules violates the octet rule?

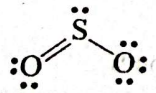
A. H_2O B. CO_2 C. SO_2 D. BH_3

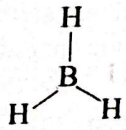
Solution

Check by drawing the correct Lewis structure

H_2O:  8 electrons on O

CO_2:  8 electrons on C

SO_2:  8 electrons on S

BH_3:  6 electrons on B

Answer: D

52. Classify the following molecules as below octet, above octet or odd number of electrons.

a. BeH_2 c. BCl_3 e. ICl_4^+

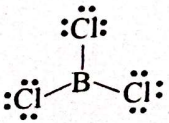
b. ICl_2^- d. ClO_2 f. NO

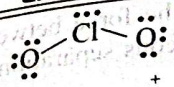
Solution

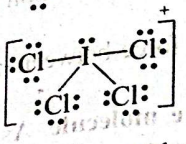
Draw the Lewis structure and classify them.

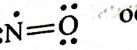
a. BeH_2: H—Be—H below octet

b. ICl_2^-: above octet

c. BCl_3:  below octet

d. ClO_2  odd number of electrons

e. ICl_4^+:  a bove octet

f. NO  odd number of electrons

Properties of Covalent Compounds

1. Covalent Compounds Exist are Molecules

These molecules occur independently.

2. Covalent Compounds Exist in the Three States of Matter.

Example

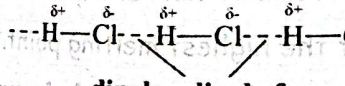
- H_2O, C_2H_5OH, C_6H_6, CCl_4 etc are liquids
- CO_2, NH_3, CH_4, C_2H_6, C_3H_8, etc are gases
- I_2, $C_{20}H_{42}$, C_6H_5COOH, $C_{25}H_{52}$, etc are solids

Note: The physical properties of covalent compounds are affected by **intermolecular forces.**

There are three types of intermolecular forces. These are

(i) **Dipole-dipole force** - is force between polar molecules

Example

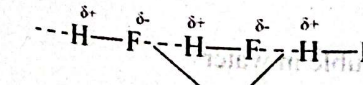


dipole - dipole force

(ii) **Hydrogen bond** is formed between polar molecules that have H - F, H - N or H - O bond.

- It is the **strongest** intermolecular force in covalent compounds

Example



Hydrogen bond

(iii) London or dispersion force is the force between the particles due to the attraction of instantaneous separation of their charge centers.

- It is the weakest intermolecular force between non-polar molecules.
- It is affected by **the size of the molecule**. As size increases, strength of London force also increases.

Example

$I_2 > Br_2 > Cl_2 > F_2$

$C_2H_{10} > C_3H_8 > C_2H_8 > C_2H_6 > CH_4$

Van der waals forces are weak forces of attraction between molecules.

Strength: Hydrogen bond > dipole-dipole force > London force

3. Generally Covalent Compounds have low Melting and Boiling points

- Polar covalent compounds have higher boiling and melting points than non-polar covalent compounds of similar size.

Reason: dipole - dipole force is stronger than London force

- As the size of non-polar covalent compounds increases, the melting and boiling points increases.

Reason: As size increases, strength of London force also increases.

Diamond exceptionally has got the highest melting point.

4. Generally polar covalent compounds are soluble in polar solvent (water) and non - polar covalent compounds are soluble in non-polar solvents

They follow the "like dissolves like" principle. Meaning polar dissolves polar and non - polar dissolves non - polar.

Example

- Sugar and ammonia are soluble in water
- oil and benzene are miscible.

5. Most covalent Compounds are Non - electrolytes.

- They are poor conductors of electricity due to lack of charge carriers.
- Some polar covalent compounds such as hydrogen chloride, hydrogen bromide, hydrogen iodide ionize in aqueous solution and conduct electricity because they produce ions that carry charge.

Graphite is an allotrope of carbon that conducts electricity

6. Covalent Compounds Generally have low Density

- Covalent compounds in liquid or gaseous state have low density due to less number of molecules per volume.

Density of solid compounds > liquid compound > gaseous compounds

7. The Reaction of Covalent Compounds is lower than Ionic Compounds.

- This is due to the lower energy required to break the bonds in covalent compounds, than ionic compounds. Dissolution of ionic compounds in water releases ions easily.

Illustrative Examples

53. Which of the following is true about covalent compounds?

A. They are molecular substances.
B. They are soluble in polar solvents
C. They shatter when hammered with external force
D. They are solids at ordinary conditions.

Answer: A

54. Which of the following is not an intermolecular force?

A. dipole-dipole force
B. London force
C. hydrogen bond
D. Covalent bond

Solution

Covalent bond is a chemical bond with in a molecule not between molecules.

55. Which of the following has the highest boiling points?

Answer: D

A. ICl B. I_2 C. Cl_2 D. N_2

Solution

ICl is a polar molecule. Due to the dipole-dipole interactions, it has the highest boiling point

56. The following table shows substance of comparable molecular masses with their dominant intermolecular forces.

Answer: A

Substance	Dominant intermolecular force
A	London force
B	Hydrogen bond
C	Dipole-dipole force

Which of the following is true about the substances?

A. "B" has the strongest intermolecular force of attraction.

B. "C" has weaker intermolecular force of attraction than "A"

C. "C" has the weakest intermolecular force of attraction.

D. "A" has higher boiling point than "B"

Solution

Hydrogen bond is the strongest intermolecular force of attraction

Answer: A

57. In which of the following substances is attraction force between the particles is the weakest?

A. Helium B. Water C. Ice D. Diamond

Solution

Noble gases are non-polar gaseous molecules whose molecules are held together with weak intermolecular force.

Answer: A

58. Define Van der waal's forces of attraction?

Solution

Van der waal's forces of attraction is the collective name for certain attractive forces between non-polar molecules. They arise because of attractive interactions between instantaneous temporary

oscillating dipoles and induced dipoles.

59. Most covalent compounds are poor conductors of electricity. Explain why?

Solution

Reason : Most covalent compounds do not produce ions that carry charge in solution or molten form.

60. Most covalent compounds are insoluble in water. **Explain why?**

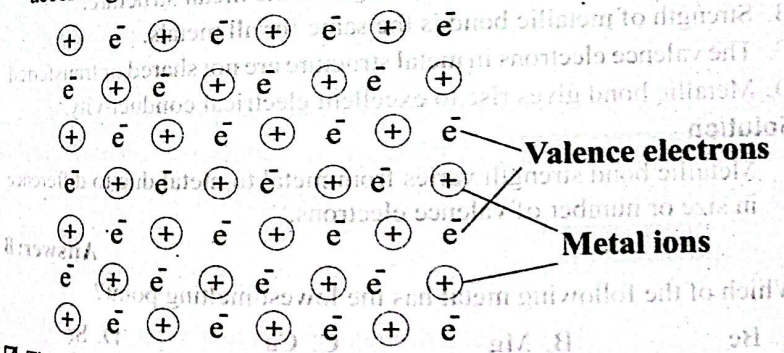
Solution

Reason: Most of them are non-polar covalent compounds that are solution in non-polar solvents.

5.4 Metallic Bonding

☑ It is electrostatic force of attraction between metal cations (kernels) and the delocalized valence electrons.

☑ Metal cations or kernels are the positive nuclei and inner shell electrons according to **Paul Drude** electron - sea model.



☑ The valence electrons of metals are also called **free electrons.**

☑ Most metals have 1, 2 or 3 valence electrons in their outer most shell

☑ The strength of metallic bond is affected by **size and number of valence electrons.**

$$\text{Metallic bond Strength} \propto \frac{\text{Number of valence electrons}}{\text{Size}}$$

☑ Metal with smaller size or more number of valence electrons has

stronger metallic bond.

Example

Li > Na > K : due to size only

Al > Mg > Na : due to size and valence electrons

Metallic Bond Accounts for the Following Physical Properties of Metals

- ☑ Metals are malleable (hammered in to sheet)
- ☑ They are ductile (drawn into wire)
- ☑ They have high melting and boiling points.
  - ➥ The melting point of metals cover a broad range from - 38.9°C (mercury) to 3410°C (tungsten).
- ☑ Metals transmit heat and electricity without undergoing chemical change because of the free electrons that carry charge.

Illustrative Examples

61. Which of the following is not true of a metallic bond?

A. Electrons are free to move throughout the metal structure.
B. Strength of metallic bond is the same for all metals.
C. The valence electrons in metal structure are not shared or transferred.
D. Metallic bond gives rise to excellent electrical conductivity.

Solution

Metallic bond strength varies from metal to metal due to difference in size or number of valence electrons.

Answer: B

62. Which of the following metal has the lowest melting point?

A. Be B. Mg C. Ca D. Sr

Solution

As size of the metal increases its melting point decreases because of decrease in metallic bond strength.

Answer: D

63. Which of the following is the most metallic element?

A. Li B. K C. Rb D. Na

Solution

As size increases, metallic character also increases and vice - versa

Answer: C

64. Explain why the melting point of calcium is higher than potassium.

Example

$_{20}Ca$: 2, 8, 8, 2

$_{19}K$: 2, 8, 8, 1

Reason : Ca has one more valence electron and smaller size that K.

65. Transition metals have particularly high melting and boiling points. Why?

Solution

Reason : They involve more valence electrons in the delocalization.

66. What makes metals to be good conductors of heat and electricity?

Solution

The mobility of delocalized electrons.

67. Identify the metals with the lowest and highest melting points.

Solution

- ➥ Lowest melting point : mercury (- 38.9°C)
- ➥ Highest melting point : Tungsten (3410°C)

68. Which member has higher boiling point ?

a. Li or Na b. K or Mg c. Al or Ca d. Na or Fe

Solution

a. Li b. Mg c. Al d. Fe

69. Which one is more metallic element in question 68?

Solution

a. Na b. K c. Ca d. Na

70. List the properties of metals that are infulenced by metallic bonding

Solution

- ➥ Malleability
- ➥ Ductility
- ➥ Conductivity
- ➥ Melting and boiling points.

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Solved Problems

71. Explain how ionization energy and electron affinity affect the formation ionic bonding.

Explanation

Elements with low ionization energy (metals) and with high electron affinity (non-metals) favor formation of ionic bonding

72. Write the Lewis dot symbols for period three elements and predict the number of bonds the atoms of elements form.

Solution

The number of unpaired dots corresponds to the number of bonds an atom of the element can form in a compound

Lewis dot symbol	Number of bonds formed
Na·	1
·Mg·	2
·Al·	3
:Si:	4
:P·	3
:S:	2
:Cl·	1
:Ar:	0

73. Arrange the following ionic compounds in decreasing order of melting point. LiF, LiI, LiBr, LiCl

Solution

The smaller the size of the cation or anion the higher is the melting point.

➥ LiF > LiCl > LiBr > LiI

74. Draw the Lewis structure of the following ionic compounds.

a. K_3N b. $SrBr_2$ c. FeF_3 d. Na_2O

Solution

a. $K_3^+[:\ddot{N}:]^{3-}$ b. $Sr^{2+}[:\ddot{Br}:]_2^-$ c. $Fe^{3+}[:\ddot{F}:]_3^-$ d. $Na_2^+[:\ddot{O}:]^{2-}$

75. Write the formula of the ionic compound formed between:

a. Ga and N b. Zn and O c. Mg and P d. Ba and F

Solution

a. GaN b. ZnO c. Mg_3P_2 d. BaF_2

76. Name the following ionic compounds

a. BaI_2 b. CsCl c. CuF_2 d. Fe_2O_3

Solution

a. Barium iodide c. Copper (II) fluoride

b. Cesium chloride d. Iron (III) oxide

77. Write three examples of ionic compounds that violate the octet rule.

Solution

Ionic compounds whose cations do not have eight electrons in their valence shells violate the octet rule.

Example

LiF : $_3Li^+ : 1s^2$: 2 electrons

FeO : $_{26}Fe^{2+} : 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6$

: [Ne] $3s^2\ 3p^6\ 3d^6 = 2 + 6 + 6 = 14$ electrons

ZnO : $_{30}Zn^{2+}$: [Ar] $3s^2\ 3p^6\ 3d^{10}\ 2 + 6 + 10 = 18$ electrons

78. What types of electrons participate in :

a. ionic bond formation?

b. Covalent bond formation?

Solution

a. Valence electrons of metals

b. Unpaired valence electrons of non-metals.

79. State the type of bonding -ionic, covalent or metallic you would expect in

a. O_3 b. $MgCl_2$ c. BrO_2 d. Fe

Solution

➥ In a and c : covalent bond

➥ In b : ionic bond

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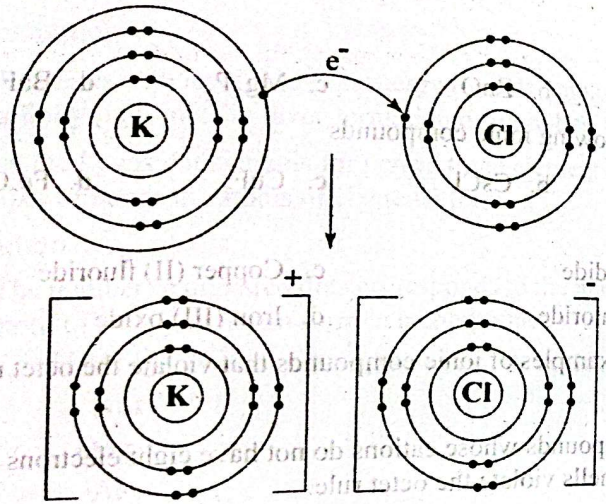
➥ In d : metallic bond

80. Explain the formation of ionic bond between potassium and chlorine using Bohr diagram.



KCl

81. The element A lies on the left side of transition elements while the element B lies on the right side of transition elements. If A and B are both univalent, predict the formula and type of the compound formed between A and B.

Solution

The compound formed is most likely ionic and the formula is $A^+ B^-$ or AB.

82. Arrange the following compounds in increasing order of ionic character.

$MgCl_2$, $CaCl_2$, $BeCl_2$, $SrCl_2$

Solution

The larger the electronegativity difference between the bonded atoms, the higher is the ionic character.

$BeCl_2 < MgCl_2 < CaCl_2 < SrCl_2$

83. Compare the bond length and bond energy of

C—O , C=O and C≡O

Solution

Bond length: C—O > C=O > C≡O

Bond energy: C≡O > C=O > C—O

84. Draw the Lewis structure and determine the number of bonding lone pair of electrons in :

a. SCl_2

b. CH_3Cl

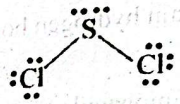
c. IF_3

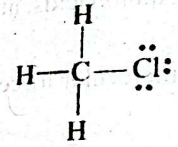
d. NO_3^-

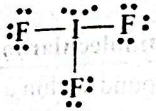
e. PH_4^+

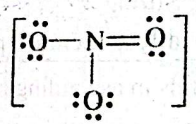
f. HNO_3

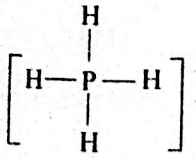
Solution

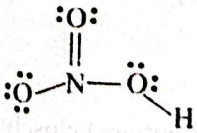
a. SCl_2:  Lp = 8, bp = 2

b. CH_3Cl:  bp = 4, Lp = 3

c. IF_3:  bp = 3, Lp = 11

d. NO_3^-:  bp = 4, Lp = 8

e. PH_4^+:  bp = 4, Lp = 0

f. HNO_3:  bp = 5, Lp = 7

85. Classify the following as polar or non- polar molecules.

a. SO_2

b. H_2Se

c. SiH_4

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d. $CHCl_3$ e. CO_2 f. CH_3NH_2

Solution

➥ a, b, d and f are polar

➥ c and e are non-polar

86. Arrange the following in increasing order of polarity

$HF, H_2S, PH_3, NH_3, HCl, H_2O$

Solution

$HF > H_2O > NH_3 > HCl > H_2S > PH_3$

87. List some group of substance that contain hydrogen bond.

Solution

Water, alcohols, carboxylic acids, amionacids, ammonia, hydrogen fluoride.

88. Explain the difference between intermolecular force and intra molecular force.

Solution

Intermolecular Force	Intramolecular force
➥ Found between molecules	➥ Found within a molecule
➥ Weak	➥ Strong
➥ affect physical property	➥ affect chemical property

89. Arrange the following organic compounds in ascending order of boiling point.

$C_{10}H_{22}, C_{15}H_{32}, C_{12}H_{26}, C_{18}H_{38}$

Solution

$C_{10}H_{22} < C_{12}H_{26} < C_{15}H_{32} < C_{18}H_{38}$

90. Explain metallic bond.

Solution

Metallic crystal consists of three dimentional closely packed assembly of metal kernels immersed in the sea of mobile valence electrons. The force of attraction between metal kernels and the mobile valence electrons is termed as metallic bond.

Supplementary Exercise

I. True or False items.

1. A chemical bond is attraction force that holds atoms or ions together in different species.
2. Ionic bond is formed by transfer or electrons from negative atom to positive atom.
3. Ionic compounds transmit elecctricity in molten state without undergoing chemical change.
4. Formation of simple cations with higher charge (+3 or above) is rare because of high ionization energy requirement.
5. Ionic compounds are hard and malleable.
6. A Lewis dot symbol is the same as a Lewis structure.
7. Intermolecular forces are weaker than intramolecular forces.
8. Covalent bond is formed by mutual sharing of electrons between atoms.
9. Higher the dipole moment of the molecule, greater is its polarity.
10. There is a polar molecule with non-polar bond.

II. Matching Items

A	B
11. Ionic bond	A. $X: + \dot{Y} \longrightarrow [X\ Y]^+$
12. Polar Covalent bond	B. $A\cdot + \cdot B \longrightarrow A:B$
13. Non-polar covalent bond	C. $E\cdot + \cdot F \longrightarrow E^+:F^-$
14. Coordinate covalent bond	D. $Z\cdot + \cdot W \longrightarrow Z^{\delta+}:W^{\delta-}$
15. Hydrogen bond	E. - - -H—Cl- - -H—Cl- - -
	F. - - -H—F- - -H—F- - -
	G. - - -O=O- - -O=O- - -

III. Multiple Choice Items.

16. When three pairs of electrons are shared ____ is formed.
A. double bond C. triple bond
B. single bond D. ionic bond

17. Which of the following is electrovalent bond?
A. Covalent bond C. Hydrogen bond
B. Ionic bond D. dative bond

18. Which of the following compound is most ionic?
A. CCl_4 B. $Sn Cl_2$ C. $Mg Cl_2$ D. $P Cl_5$

19. Which of the following bond is most polar ?
A. C—H B. S—H C. N—H D. O—H

20. The pair of elements most likely to form ionic compound is
A. Al and Cl_2 B. O_2 and Cl_2 C. Cl_2 and I_2 D. Cs and Cl_2

21. Which of the following process involves cleavage of hydrogen bonds?
A. Sublimation of dry ice C. Dissociation of H_2 molecules
B. Melting of ice D. Condensation of ammonia

22. Which of the following will have maximum covalent character?
A. NaCl B. NaI C. NaBr D. NaF

23. The electronic configuration of metal M is $1s^2\ 2s^2\ 2p^6\ 3s^2$. The formula of its oxide will be
A. MO B. M_2O C. M_2O_3 D. MO_2

24. Which of the following does not transmit electricity?
A. Salt Solution C. Graphite
B. Gold D. Sugar solution

25. Which of the following molecules has no lone pair of electrons on the centeral atom?
A. H_2O B. SO_2 C. SO_3 D. OF_2

26. When viewed structurally, ionic compounds generally are:
A. Single molecule
B. Gaint lattices
C. Ions surrounded by a sea of electrons
D. Gaint molecules

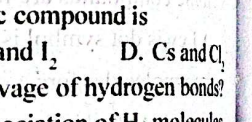
27. An element with atomic number 19 would most likely combine with an element with atomic number _____ to ionic compound.
A. 10 B. 9 C. 11 D. 12

28. Which of the following contains double bond ?
A. N_2 B. NH_2 C. Cl_2 D. SO_3

29. Two atoms A and B have electron configuration of $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$ and $1s^2\ 2s^2\ 2p^4$ respectively. Which one is true?
A. An ionic compound of formula AB is formed
B. A ionic compound of formula A_2B_2 is formed
C. An ionic compound of formula AB_2 is formed
D. A covalent compound of formula A_2B is formed

30. In which of the following compounds are all ions isoelectronic with the argon atom?
A. KBr B. Al_2O_3 C. $CaCl_2$ D. NaF

31. Given the Lewis structure of chloro methane as follows:



How many lone pair of electrons are there?
A. 3 B. 7 C. 5 D. 4

32. What do we call the attractive force that operates between molecules?
A. Intra molecular force C. Molecuar binding force
B. Inter molecular force D. Hydrogen bond

33. Which of the following does not contain any dative bond?
A. H_3O^+ B. $Na BF_4$ C. NH_4Cl D. HF_2^-

34. Which among the following has highest boiling point ?
A. He B. CH_4 C. CO D. Ne

35. Which of the following molecules has the shortest nitrogen to nitrogen bond?
A. N_2 B. N_2H_4 C. N_2O D. N_2O_4

36. Which of the following is the most metallic element?
A. Al B. B C. C D. Si

37. All of the following are intramolecular forces. EXCEPT
A. Ionic bond C. Covalent bond
B. Hydrogen bond D. Metallic bond

38. Which of the following covalent compounds obeys the octet rule?

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A. ICl_3 B. PCl_3 C. BrF_3 D. ClF_3

39. Which of the following compounds has only ionic bond?
A. NaF B. BrF C. NH_4F D. NaCN

40. All of the following has coordinate covalent bond. EXCEPT
A. CO B. Al_2Cl_6 C. NH_4NO_3 D. NH_3

IV. Short Answer Items

41. _____ shows the number of valence electrons possessed by an atom of given element.

42. The bond length of A - B, A - C, A - D and A - E is $0.5A^0$, $0.78A^0$, 1.6^0A and $0.42A^0$ respectively. Arrange these bonds in increasing order of bond strength.

43. Ions are formed by _____ or _____ of electrons.

44. In general, an ionic bond is formed between _____ and _____

45. A non-polar covalent bond is formed when the electronegativity difference between the atoms is in the range of _____.

46. An ionic bond can be cleaved by ____ energy.

47. How atoms in a compound or molecule obey the octet rule?

48. Explain the "like dissolves like" principle.

49. Which member in each pair would you expect to exhibit hydrogen bonding?
a. CH_3OH or CH_3OCH_3
b. CH_3COOCH_3 or CH_3COOH

50. Arrange the following in increasing order of boiling point.
N_2, Br_2, Mg, K.

Answer key to Supplementary Exercise on unit - 5

Part I

1. True
2. False
3. False
4. True
5. False
6. False
7. True
8. True
9. True
10. False

Part II

11. C
12. D
13. B
14. A
15. F

Part III

16. C
17. B
18. C
19. D
20. D
21. B
22. B
23. A
24. D
25. C
26. B
27. B
28. D
29. A
30. C
31. B
32. B
33. D
34. C
35. A
36. A
37. B
38. B
39. A
40. D

Part IV

41. A Lewis dot symbol

42. A - E < A - B < A - C < A - D

43. Losing or gaining

44. Metals and non-metals

45. 0.0 - 0.5

46. Heat

47. By gaining losing or sharing of valence electrons.

48. Polar dissolves polar and non - polar dissolves non- polar.

49.
a. CH_3OH
b. CH_3COOH

50. $N_2 < Br_2 < K < Mg$

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UNIT 6

Chemical Reactions and Stoichiometry

Introduction

One of the concern of chemistry is studying the changes in which matter undergoes. These changes are the physical and chemical changes.

Physical Change is a change that results no new substance.

- ➥ It is a change in physical property
- ➥ It is easily reversed by physical means.
- ➥ It doesn't alter the composition of a substance
- ➥ The energy change is not necessary.

Table : The Three Types of Physical Changes and their Examples

Types of physical changes	Example
Change of state	➥ Melting (solid→liquid) ➥ Freezing(liquid→solid) ➥ Evaporation(liquid→gas) ➥ Condensation(gas→liquid) ➥ Sublimation(solid→gas) ➥ Deposition(gas→solid)
Mechanical change (Change in shape or size)	➥ Powdering of chalk ➥ Tearing of paper ➥ Cutting of wood,..etc..
Dissolution	➥ Dissolution of salt ➥ Dissolution of sugar ➥ Mixing ethanol and water

Chemical Change: - is a change that results new substance.

- ➥ It is accompanied by energy changes.
- ➥ It is generally difficult to reverse.
- ➥ It alters the composition of the substance.
- ➥ It involves the chemical property of a substance.

Example:- Rusting of iron
Fermentation
Respiration
Photosynthesis
Reaction of sodium with water
Decaying of food. etc...

Illustrative Examples

1. Which of the following changes is an example of a chemical change?
   A. Decomposition of copper carbonate
   B. Melting of fat
   C. Evaporation of water
   D. Dissolution of salt in water

   Solution:
   Decomposition is a chemical reaction that results new substance. Thus, it is a chemical change

   Answer: A

2. Among the following which one is change of state?
   A. Sublimation
   B. Breaking of glass.
   C. Dissolution of sugar
   D. Powdering of chalk

   Answer: A

3. Identify the wrong statement about a chemical change.
   A. New substance is formed
   B. Alters the composition of a substance
   C. Energy is absorbed or released

D. It is easily reversed

Solution:

Chemical change is difficult to reverse.

Answer: D

4. What are the three types of physical changes?

Solution:

- Change of state
- Mechanical change
- Dissolution

5. Classify the following as chemical or physical change.

a. Burning of charcoal
b. Sublimation of iodine
c. Condensation of water vapour
d. Heating of sugar
e. Fermentation of Tella
f. Souring of milk
g. Formation of cloud
h. Mixing of mud and water

Solution:

- a,d,e and f are chemical changes
- b,c,g and h are physical changes

6.1 Chemical Reactions

A chemical reaction is a chemical change.

- ☑ It transforms reactants into products.
- ☑ **Reactants** are raw materials or starting materials, usually written to the **left** side of the arrow.
- ☑ **Products** are the new substances formed during chemical reaction. They are usually written to the **right** side of the arrow.

Reactant →Product

- ☑ The arrow (→) implies "results" or "gives" or "yields" or "produces".
- ☑ **Chemical equation** uses chemical symbols and formula to show what happens during a chemical reaction.
- ☑ It is a short hand representation of a chemical reaction by means of chemical symbols and formula.

Writing Chemical Equations

- ☑ A correct chemical equation shows what changes take place.
- ☑ It also shows the relative amounts of the various elements and compounds that take part in these changes.
- ☑ It usually indicates the physical states of substances involved with small letters in parentheses, (s - solid, l - liquid and g - gas)

Rules

To write the Correct Chemical Equation

step 1: Write the **Word equation** of the chemical reaction.

step 2: Transform the word equation into **formula or symbolic equation** and write the physical states of substance involved.

step 3: Balance the chemical equation

Example: Consider the reaction between hydrogen and oxygen to yield water.

step 1: Hydrogen + Oxygen →Water

step 2: $H_2(g) + O_2(g) \rightarrow H_2O(l)$

step 3: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

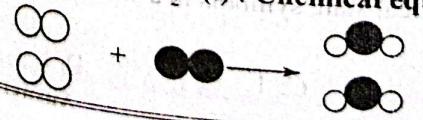
Physical state symbols in a chemical equation

- (g) -gas
- (ℓ) - liquid
- (s) -solid
- (aq) -water solution of substance

Table : 6.2 Representation of Chemical Equation:

Hydrogen + Oxygen → water : **word equation**

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$: **Chemical equation**



: Model equation

Table : 6.3 Interpretation of a Chemical Equation

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

Two molecules + one molecule → 2 molecules

2 moles + one mole → 2 moles

2(2g) + 1(32g) → 2(18g)

36g reactants 36g product

The names of reactants and products represent **the qualitative meaning** of a chemical equation.

The amounts of reactants and products tell us the **quantitative meaning** of a chemical equation.

Illustrative Examples

6. Which of the following represents the correct chemical equation for burning of magnesium to yield magnesium oxide?

A. $Mg(s) + O(g) \rightarrow MgO(s)$ C. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

B. $Mg(s) + O_2(g) \rightarrow MgO(s)$ D. $2Mg(s) + O_3(g) \rightarrow 3MgO(s)$

Solution:

A chemical equation must be represented by correct symbols and formula. The seven diatomic molecules occur in a chemical equation in molecular form not in an atomic form.

(H_2, O_2, N_2, F_2, Cl_2, Br_2 and I_2)

The formula of magnesium oxide is MgO. **Answer: C**

7. Chemical equations are used to represent____

A. Physical change C. Compounds

B. Elements D. Chemical reactions.

Answer: D

8. All of the following are physical state symbols in a chemical equation. **EXCEPT**

A. g B. p C. l D. s

Answer: B

9. Define

a. Reactants b. Products

c. Chemical equation d. Chemical reaction

Solution:

a. Reactants are starting substances in a chemical reaction

b. Products are the new substance formed in chemical reactions

c. Chemical equation is short hand representation of a chemical reaction.

d. Chemical reaction is the process by which one or more substance are changed into one or more different substances.

10. Identify the reactants and products in the chemical equations below.

a. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

b. $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$

c. $CaCO_3(g) \rightarrow CaO(s) + CO_2(g)$

Solution:

	Reactions	Products
a.	N_2 and H_2	NH_3
b.	C_2H_6 and O_2	CO_2 and H_2O
c.	$CaCO_3$	CaO and Co_2

11. How could you represent in short, the formation of carbon dioxide and water from burning of a acetylene gas?

Solution:

$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$

12. What is the meaning of the plus sign (+) in the reactants side?

Solution:

Plus (+) means reacts or combines.

6.2 Balancing Chemical Equation

- Balanced chemical equations have the same kind and number of atoms on each side.
- **Atoms, charge** and **mass** are conserved in a balanced chemical equation.
- **Coefficients**, not **subscripts** may be changed to balance an equation. Changing a subscript changes the substance that is taking part in the reaction.

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Some of the Methods Used to Balance Chemical Equation are:

1. The inspection method
2. The least common multiple method
3. The algebraic method
4. The oxidation number method
5. The ion-electron method

The last two will be discussed in higher level. They are used to balance oxidation-reduction reaction.

1. The Inspection Method

- It is also called **trial and error method.**
- It is mainly used to balance simple chemical equations.
- Place the smallest whole number coefficients for the unbalance atoms.

Example:

$H_2 + O_2 \rightarrow H_2O$ (unbalanced)

Here, hydrogen is balanced but not oxygen. Place 2 as a coefficient for H_2O. $H_2 + O_2 \rightarrow 2H_2O$

Now, oxygen is balanced but not hydrogen. Then, write 2 as a coefficient for H_2 to balance it. $2H_2 + O_2 \rightarrow 2H_2O$ (Balanced)

Lastly check weather the number of atoms and masses are conserved or not.

Reactants	Product
$H : 2 \times 2 = 4$	$H : 2 \times 2 = 4$
$O : 1 \times 2 = 2$	$O : 2 \times 1 = 2$
Total atoms $= 2 + 4 = 6$	Total atoms $= 4 + 2 = 6$
Mass: $4g + 32g = 36g$	Mass $= 2 \times 18g = 36g$

2. The Least Common Multiple Method (LCM method)

This method is also used to balance simple chemical equations.

To balance by LCM method:

- Place total valence number (subscript times valence number) on each atom above it.
- Find the LCM of each total valence number and place it on the arrow
- Divide the LCM by each total valence number to obtain the coefficients of reactants and products.
- Check the conservation of atoms and masses.

Example:

$Fe + O_2 \rightarrow Fe_2O_3$ (unbalanced)

Here, valence number of Fe = 3 and O = 2

$$\overset{3}{Fe} + \overset{4}{O_2} \xrightarrow{12} \overset{6}{Fe_2}\overset{6}{O_3}$$

$12 \div 3 = 4$

$12 \div 4 = 3$

$12 \div 6 = 2$

$4Fe + 3O_2 \longrightarrow 2Fe_2O_3$

Check: atoms and masses are conserved

3. Algebraic Method (The a,b,c, . . . method)

- The method need the knowledge of mathematics
- Place small letters a,b,c, . . as coefficients and find their values by assuming the value of 'a' to be one.
- It is used to balance even complicated chemical equations.

Example:

$N_2 + H_2 \rightarrow NH_3$ (taken from the previous extreme chemistry)

Solution:

Place small letter a,b and c as coefficients of reactants and product.

$aN_2 + bH_2 \rightarrow cNH_3$

From mathematical equation to conserve atoms

$N : 2a = c$ $\qquad$ $H : 2a = 3c$

Two equation but three variables. It is impossible to find the values of the variables.

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Thus, assume a = 1

➥ $2a = c$

$c = 2 \times 1 = 2$

➥ $2b = 3c$

$2b = 3 \times 2 = 6 \Rightarrow b = 3$

Substitute the numbers in the places of the letters.

➥ $N_2 + 3H_2 \rightarrow 2NH_3$

Atoms, masses and charges are conserved.

Illustrative Examples

13. A balanced chemical equation gives the following information. **EXCEPT**

A. The nature of substance involved in the reaction.
B. The amount of substance involved in the reaction.
C. The physical states of substance involved in the reaction.
D. The rate of chemical reaction.

Solution:

From balanced chemical reaction we can understand:

➥ Nature of substance involved.
➥ Physical states of substance.
➥ Amount of substances involved

Answer: D

14. In every balanced chemical equation each side of the equation has the same number of:

A. Coefficients B. Moles C. Molecule D. Atoms

Solution:

Number of atoms, mass and charge are equal on both sides of the chemical equation, Coefficients, moles and number of molecules are not necessarily conserved.

Answer: D

15. When the reaction below is correctly balanced, what would be coefficient of HNO_3? $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O$

A. 6 B. 10 C. 8 D. 4

Solution:

Since the reaction is more complex, let us balance it by algebraic

method.

$a\,Cu + b\,HNO_3 \rightarrow c\,Cu(NO_3)_2 + d\,NO + e\,H_2O$

$Cu : a = c \ldots (1)$

$H : b = 2e \ldots (2)$

$N : b = 2c + d \ldots (3)$

$O : 3b = 6c + d + e \ldots (4)$

Assume $a = 1$

$\Rightarrow c = 1$

$b = 2e = 2 \times \frac{4}{3} = \frac{8}{3}$

$d = 2e - 2 = 2 \times \frac{4}{3} - 2 = \frac{2}{3}$

Relate (2) and (3)

$2e = 2c + d$

$2e = 2 \times 1 + d$

$2e = 2 + d$

$d = 2e - 2 \ldots (5)$

Relate (4) and (5)

$3b = 6 \times 1 + 2e - 2 + e$

but, $b = 2e$

$3(2e) = 6 + 3e - 2$

$6e - 3e = 4$

$3e = 4$

$e = \frac{4}{3}$

$\therefore a = 1 \quad d = \frac{2}{3}$

$b = \frac{8}{3} \quad e = \frac{4}{3}$

$c = 1$

} multiply by 3 to convert into positive integers

$a = 3 \times 1 = 3$ $\quad b = 3 \times \frac{8}{3} = 8$ $\quad c = 3 \times 1 = 3$

$d = 3 \times \frac{2}{3} = 2$ $\quad e = 3 \times \frac{4}{3} = 4$

$\Rightarrow 3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

Answer: C

16. Given: $Al_2(SO_4)_3 + ZnCl_2 \rightarrow AlCl_3 + ZnSO_4$, What is the sum of the coefficients of reactants and products?

Solution:

This is simple chemical reaction. We may use LCM method to

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balance it.

$$\overset{6}{Al_2}(\overset{6}{SO_4})_3 + \overset{2}{Zn}\overset{2}{Cl_2} \longrightarrow \overset{3}{Al}\overset{3}{Cl_3} + \overset{2}{Zn}\overset{2}{SO_4}$$

$$Al_2(SO_4)_3 + 3ZnCl_2 \longrightarrow 2AlCl_3 + 3ZnSO_4$$

Thus, sum of coefficients: $1 + 3 + 2 + 3 = 9$

17. What is the coefficient ratio of products after balancing the chemical equation below?

$$C_{10}H_{22} + O_2 \rightarrow CO_2 + H_2O$$

Solution:

The balanced equation becomes

$$2C_{10}H_{22} + 31O_2 \rightarrow 20CO_2 + 22H_2O$$

Thus, the coefficient ratio of products CO_2 to H_2O is 10:11.

18. Balance the following chemical equation by any of the methods you like,

a. $Fe + Cl_2 \rightarrow FeCl_3$
b. $HNO_3 + SO_2 + H_2O \rightarrow H_2SO_4 + NO$
c. $HCl + K_2Cr_2O_7 \rightarrow KCl + CrCl_3 + Cl_2 + H_2O$
d. $CuO + NH_3 \rightarrow N_2 + Cu + H_2O$

Solution:

a. $2Fe + 3Cl_2 \rightarrow 2FeCl_3$
b. $2HNO_3 + 3SO_2 + 2H_2O \rightarrow 3H_2SO_4 + 2NO$
c. $14HCl + K_2Cr_2O_7 \rightarrow 2KCl + 2CrCl_3 + 3Cl_2 + 7H_2O$
d. $3CuO + 2NH_3 \rightarrow N_2 + 3Cu + 3H_2O$

19. Write balanced chemical equation for the reaction between nitrogen pent oxide and water to produce nitric acid.

Solution:

step 1: Write the words equation

Nitrogen pent oxide + Water → Nitric acid

step 2: Write formula equation

$N_2O_5 + H_2O \rightarrow HNO_3$

step 3: Balance the equation

$N_2O_5 + H_2O \rightarrow 2HNO_3$

20. Change the following chemical equations into word equations.

a. $4NO_2 + O_2 \rightarrow 2N_2O_5$
b. $2NaHCO_3 + Na_2CO_3 + CO_2 + H_2O$

Solution:

a. Nitrogen dioxide + oxygen → dinitrogen pent oxide
b. Sodium hydrogen carbonate → sodium carbonate + carbon dioxide + water

21. What is the difference between chemical reaction and chemical equation?

Solution:

A chemical reaction is a process of converting reactants into products where as a chemical equation is a short hand expression of a chemical reaction.

22. We change only coefficients to balance a chemical equation but not subscripts. Why?

Solution:

Changing coefficients does not change the identify of substances but when we change subscripts the identity of substance change even we may get unknown formula.

6.3 Types of Chemical Reactions

There are four types of chemical reactions. These are:

1. Combination reaction
2. Decomposition reaction
3. Single displacement reaction
4. Double displacement reaction

1. Combination Reaction

- It is also called **addition reaction or synthesis**
- Two or more reactants produce only one product.

$$X + Y \rightarrow XY$$

Example:

$$BaO + H_2O \rightarrow Ba(OH)_2$$
$$2H_2 + O_2 \rightarrow 2H_2O$$

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2. Decomposition Reaction

- It is also known as **analysis**
- One reactant is converted into two or more products.

$AB \rightarrow A + B$

Example:

$2Ag_2O \rightarrow 4Ag + O_2$

$2KClO_3 \rightarrow 2KCl + 3O_2$

$2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$

There are different types of decomposition reactions

(i) Decomposition of Metallic Carbonates

Metallic carbonates except that of sodium and potassium, decompose when heated.

Metallic carbonate → Metallic oxide + carbon dioxide

Example:

$CaCO_3 \rightarrow CaO + CO_2$

(ii) Decomposition of Metallic Nitrates

- Alkali metals nitrates → Metal nitrite + Oxygen

Example:

$2NaNO_3 \rightarrow 2NaNO_2 + O_2$

Other metals nitrates → Metallic oxide + Nitrogen dioxide + oxygen

Example:

$2Zn(NO_3)_2 \xrightarrow{\Delta} 2ZnO + 4NO_2 + O_2$

3. Single displacement Reaction

- It is also called **Substitution** or **replacement reaction**
- **More** reactive element displaces less reactive element.

$X + YZ \rightarrow XZ + Y$

if X is metallic

$X + YZ \rightarrow YZ + Z$ if

X *is non - metal*

Example:

$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

Zn is more reactive than Cu

$Cu + ZnSO_4 \rightarrow$ No reaction

⇒ Cu is less reactive than Zn

$Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$

Cl_2 is more reactive than Br_2

Reactivity of halogens : $F_2 > Cl_2 > Br_2 > I_2$

Note: i. Active metals displace hydrogen from dilute acids.

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

ii. Very active metals displace hydrogen gas from water.

$2K + 2H_2O \rightarrow 2KOH + H_2$

$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$

4. Double Displacement Reaction

It is also called **Methametical reaction** or **double replacement reaction.**

$AB + CD \rightarrow AD + CB$

Example:

$NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$

- When two soluble salts react, one soluble and one insoluble salt are produced.
- For soluble salt, we write **(aq)** in bracket and for insoluble salt **(s)** or **(↓)**

Example:

$K_2SO_4(aq) + BaI_2(aq) \rightarrow BaSO_4(s) + 2KI(aq)$

Or

$K_2SO_4(aq) + BaI_2(aq) \rightarrow BaSO_4\downarrow + 2KI(aq)$

The reaction between acid and base (neutralization reaction) is an example of double displacement reaction.

All neutralization reactions are double displacement reactions but all double displacement reactions are not necessarily neutralization reactions.

Example:

$KOH(aq) + HI(aq) \rightarrow KI(aq) + H_2O(1)$

The reaction between carbonates and dilute acids is an example of double displacement reaction.

Example:

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(1) + CO_2\uparrow$

➥ ↑- Implies gas or released.

Illustrative Examples

23. Which of the following is an example of combination reaction?

A. $BaCl_2 + NaSO_4 \rightarrow BaSO_4 + H_2$
B. $Na_2O + H_2O \rightarrow 2NaOH$
C. $Zn + 2HCl \rightarrow ZnCl_2 + H_2$
D. $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$

Solution:

Two or more reactants give one product in combination reaction.

Answer: B

24. The type of reaction in which two or more products are formed from a single reactant is

A. Combination
B. Decomposition
C. Double displacement
D. Single displacement

Answer: B

25. Identify the double displacement reaction among the following.

A. An acid reacting with metal hydroxide
B. Heating metal carbonates
C. An active metal reacting with dilute acids
D. A metal reacting with a non-metal

Solution:

Acid + base → salt + water : double displacement reaction

Metal carbonate → Metal oxide + carbon dioxide : decomposition reaction

Active metal + acid → salt + hydrogen: single displacement reaction

Metal + non - metal → salt : combination reaction

Answer: A

26. The reaction between chlorine gas and potassium iodide can be classified as:

A. Single displacement reaction
B. Double displacement reaction
C. Decomposition reaction
D. Combination reaction

Solution:

$Cl_2 + 2KI \rightarrow 2KCl + I_2$

Answer: A

27. Classify the following reactions as combination, decomposition, single displacement or double displacement reaction.

a. $4HNO_3 \rightarrow 4NO_2 + 2H_2O + O_2$
b. $2Fe + 3Cl_2 \rightarrow 2FeCl_3$
c. $Pb(NO_3)_2 + Na_2SO_4 \rightarrow PbSO_4 + 2NaNO_3$
d. $Br_2 + 2LiI \rightarrow 2LiBr + I_2$
e. $Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 + H_2O$
f. $Na_2CO_3.7H_2O \rightarrow Na_2CO_3 + 7H_2O$
g. $2SO_2 + O_2 \rightarrow 2SO_3$
h. $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

Solution:

➥ a and f are decomposition reactions
➥ b and g are combination reactions
➥ c and e are double displacement reactions
➥ d and h are single displacement reactions

28. Given the reactivity of metals: Mg > Zn > Fe > Cu. Complete the following reactions.

a. $Fe + CuSO_4 \rightarrow$
b. $Zn + FeSO_4 \rightarrow$
c. $Cu + ZnSO_4 \rightarrow$
d. $Mg + FeSO_4 \rightarrow$

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Solution:

More reactive metal displaces less reactive metal but not the reverse.

a. $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$

b. $Zn + FeSO_4 \rightarrow ZnSO_4 + Fe$

c. $Cu + ZnSO_4 \rightarrow$ no reaction: because Cu is less reactive than Zn.

d. $Mg + FeSO_4 \rightarrow MgSO_4 + Fe$

6.4 Oxidation and Reduction Reactions

1. Oxidation Number or Oxidation State

The number of electrons lost or gained when an atom combines with another atom is known as **Oxidation number** or **oxidation state.**

Rules to Determine Oxidation Number

1. The oxidation number of **free atoms and molecules of the same atom is zero.**

   Example:

   Mg, S, Cl, P_4, S_8...etc... have zero oxidation number.

2. The oxidation number of **ions** is the charges on the ions.

   Example:

   $Na^+ = +1$, $S^{2-} = -2$, $P^{3-} = -3$, $NH^+_4 = +1$ etc....

3. The oxidation number of **alkali metals** (Li, Na, K, Rb, Cs) in compounds is +1

   Example:

   NaCl, Na = +1

4. The oxidation number of **alkaline earth metals** (Be, Mg, Ca, Sr, Ba) in compounds is +2

   Example:

   $CaCl_2$, Ca = +2

5. The oxidation number of **aluminium** in compounds is +3.

Example:

AlF_3, Al = +3

6. The oxidation number of oxygen in most compounds is -2. However, in:
   - Peroxides is -1: H_2O_2, O = -1
   - Super oxides is -1/2: KO_2, O = -1/2
   - Compound with fluorine is +2: OF_2, O = +2

7. The oxidation state of **halogens** (F, Cl, Br, I) in binary compounds is **-1.**

   Example:

   NaCl : Cl = -1

   AlF_3 : F = -1

8. The oxidation state of **hydrogen** in most compounds is **+1** but in metal hydrides is **-1**

   Example:

   H_2O : H = +1

   NaH : H = -1

9. The sum of oxidation number of elements in a neutral compounds is **zero.**

   Example:

   In $KMnO_4$: oxidation number of : K + Mn + O = 0.

10. The sum of oxidation number of elements in a polyatomic ion is **the charge on the ion.**

    Example:

    SO_4^{2-} : Oxidation number of : S + O = -2

Illustrative Examples

29. The oxidation number of Na , O_2, Cl^- and NH^+_4 respectively are

A. +1, -4, -1, +1

(B) 0, 0, -1, +1

C. 0, -2, -1, +5

D. +1, -2, -1, +5

Solution:

Na is free atom, thus its oxidation number is zero.

O_2 is free molecule, thus its oxidation number is zero

Cl^- is an ion, it has -1 oxidation state

NH^+_4 is an ion, it has +1 oxidation state

Answer: B

30. What is the oxidation number of sulphur in the molecule of $H_2S_2O_7$?

A. +2 B. +3 C. +5 D. +6

Solution:

Let the oxidation number of sulphur (S) be x. $H^{+1}_2 S^{x}_2 O^{-2}_7$

$2(+1)+2(x)+7(-2)=0$ because it is neutral

$2+2x-14=0$

$2x-12=0$

$\frac{2x}{2}=\frac{+12}{2}=+6$, Thus the oxidation number of S in $H_2S_2O_7$ is +6

Answer: D

31. The oxidation number of carbon in the compound $NaHCO_3$ is

A. -2 B. -4 C. +6 D. +4

Solution:

$\overset{+1}{Na}\overset{+1}{H}\overset{x}{C}O^{-2}_3$

$1\times(+1)+1\times(+1)+1\times(x)+3(-3)=0$

$1+1+x-6=0$

$2-6+x=0$

$x=+4$, Thus the oxidation number of carbon in $NaHCO_3$ is +4.

Answer: D

32. Which of the following is not true about oxidation number?

A. The oxidation number of hydrogen in CaH_2 is -1

B. The oxidation number of hydrogen in peroxide is +1

C. The sum of oxidation number of all atoms in a compound is positive or negative numbers.

D. The oxidation number of a neutral atom and molecule is zero.

Answer: C

33. Determine the oxidation state of :

a. Fe in Fe_3O_4

b. P in $K_4P_2O_7$

c. Fe in $Na_4(Fe(CN)_6)$

d. S in $S_4O^{2-}_6$

e. N in NH^+_4

f. V in VO^+_2

g. Cr in $Cr_2O^{2-}_7$

h. F in NaF

Solution:

a. +8/3 c. +2 e. -3 g. +6

b. +5 d. +5/2 f. +5 h. -1

34. Determine the oxidation number of nitrogen in the five oxides of nitrogen.

Solution:

N_2O: $N^x_2O^{-2}$: $2x+(-2)=0 \Rightarrow x=+1$

NO: N^xO^{-2}: $x+(-2)=0 \Rightarrow x=+2$

NO_2: $N^xO^{-2}_2$: $2x+2(-2)=0 \Rightarrow x=+4$

N_2O_3: $N^x_2O^{-2}_3$: $2x+3(-2)=0 \Rightarrow =+3$

N_2O_5: $N^x_2O^{-2}_5$: $2x+5(-2)=0 \Rightarrow x=+5$

2. Oxidation and Reduction

Early Definition of Oxidation

- **Oxidation** is addition of oxygen to a substance or is the removal of hydrogen from a substance.

Modern Definition of Oxidation

- **Oxidation** is loss of electrons. Relate with "oil" "to" remember oxidation "Oil" : O - Oxidation, i - is, l - loss.
- It is an increase in oxidation number.
- **Oxidation reaction** is the half - reaction that represents the loss of electrons.

Example:

$Fe \rightarrow Fe^{s+} + 3e^-$

$2Cl^- \rightarrow Cl_2 + 2e^-$

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Early Definition of Reduction

Reduction is the addition of hydrogen to a substance or removal of oxygen from a substance.

Modern Definition of Reduction

- **Reduction** is gain of electrons. Use **"Rig"** to remember it easily.
- **Rig:** **R**- Reduction, **i**- is, **g**- gain.
- It is decrease in oxidation number
- **Reduction reaction** is the half-reaction that represents the gain of electrons.

Example:

$$Cu^{2+} + 2e \rightarrow Cu$$

$$S + 2e^- \rightarrow S^{2-}$$

In a given reaction, if there is oxidation there must be reduction and vice versa.

3. Oxidizing and Reducing Agents

Oxidizing agent is a species being reduced but causes oxidation

- It gains electrons
- It exhibits decrease in oxidation number.

Reducing agent is a substance being oxidized but causes reduction.

- It loses electrons
- It exhibits increases in oxidation numbers.

Example:

Identify the oxidizing and reducing agents in the reaction given below

$$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$$

Solution:

Both are found in the reactants side.

- **Zn** is reducing agent because its oxidation number increased from zero to +2.
- **Cu^{2+}** is oxidizing agent since it gained two electrons or its oxidation number reduced from +2 to zero.

4. Oxidation - reduction or Redox Reaction

- It is a type of reaction that involves transfer of electrons.
- It is the combination of oxidation and reduction half - reactions.

Example:

- $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$
- $Cu + 2Ag^+ \rightarrow Cu^{2+} + Ag$
- $3H_2S + 2HNO_3 \rightarrow 3S + 2NO + 4H_2O$

Disproportionation reaction is a reaction in which the same substance is reduced and oxidized simultaneously.

Example:

$$Cl_2 + H_2O \rightarrow HCl + HClO$$

Here, Cl_2 is oxidized and reduced at the same time.

Note :

- *All disproportionation reactions are redox reactions but the reverse is not true.*
- *All reactions involving free atoms or molecules are redox reactions.*
- *Non - redox reactions are those reactions that do not show include and decrease in oxidation number.*
- *All double displacement reactions are non - redox reaction*

$$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$$

Balancing Redox Reaction

We may balance redox reaction by using inspection, LCM or algebraic method. However, the best method used to balance redox reaction is **Oxidation - number method.**

Steps

step 1: Identify the oxidized and reduced atoms by assigning oxidation number.

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Example:

$$\overset{+1}{H_2}\overset{-2}{S} + \overset{+1}{H}\overset{+5}{N}\overset{-2}{O_3} \longrightarrow \overset{0}{S} + \overset{+2}{N}\overset{-2}{O} + \overset{+1}{H_2}\overset{-2}{O}$$

Here, S oxidized and N is reduced

step 2: Balance the reduced and oxidized atoms.

$H_2S + HNO_2 \rightarrow S + NO + H_2O$ (already balanced)

step 3: Determine the total electrons lost and gained and connect with line and write the electrons lost and gained.

$$\overset{-2}{H_2S} + \overset{+5}{HNO_3} \longrightarrow \overset{0}{S} + \overset{+2}{NO} + H_2O$$

3e⁻ gained

2e⁻ lost

step 4: Balance the electron lost and gained by finding their LCM

LCM of 2 and 3 is 6

$\Rightarrow 6 \div 2 = 3$ and $6 \div 3 = 2$

$\Rightarrow$ Write 3 as coefficient for H_2S and S and 2 for HNO_3 and NO

$$3H_2S + 2HNO_3 \longrightarrow 3S + 2NO + H_2O$$

step 5: Balance the rest of the atoms by inspection method.

$3H_2S + 2HNO_3 \rightarrow 3S + 2NO + H_2O$

On the reactants side we have 8H - atoms.

Thus, write 4 as coefficient for H_2O

$3H_2S + 2HNO_3 \rightarrow 3S + 2NO + 4H_2O$

step 6: Check whether the reaction is balanced or not .

$3H_2S + 2HNO_3 \rightarrow 3S + 2NO + 4H_2O$

Reactants	Products
$H: 3 \times 2 + 2 = 8$	$H: 4 \times 2 = 8$
$S: 3 \times 1 = 3$	$S: 3 \times 1 = 3$
$N: 2 \times 1 = 2$	$N: 2 \times 1 = 2$
$O: 2 \times 3 = 6$	$O: 2 \times 1 + 4 \times 1 = 6$

Illustrative Examples

35. What happens to a reducing agent in an oxidation - reduction reaction?

A. It is reduced as it gains electron|s|
B. It is reduced as it loses electron|s|
C. It is oxidized as it gains electron|s|
D. It is oxidized as it loses electron|s|

Solution:

Reducing agent:

- Causes reduction
- It self oxidized
- Loses electrons

Answer: D

36. Which of the following is an oxidation - reduction reaction ?

A. $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
B. $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$
C. $CuO + H_2 \rightarrow Cu + H_2O$
D. $HCl + H_2O \rightarrow H_3O^+ + Cl^-$

Solution:

A reaction with free atom or molecule is a redox reaction.

Answer: C

37. Which of the following statements about the reaction given below is true?

$14HCl + K_2Cr_2O_7 \rightarrow 2KCl + 2CrCl_3 + 3Cl_2 + 7H_2O$

A. HCl is reducing agent
B. Hydrogen is oxidized
C. Chromium is oxidized
D. Potassium is reduced

Solution:

$$\overset{-1}{H}\overset{-1}{Cl} \longrightarrow \overset{0}{Cl}$$

$\Rightarrow$ Chlorine is oxidized

$\Rightarrow$ HCl is reducing agent

$$K_2\overset{+6}{Cr_2}O_7 \longrightarrow 2\overset{+3}{Cr}Cl_3$$

$\Rightarrow$ Cr is reduced

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$K_2Cr_2O_7$ is oxidizing agent.

Answer: A

38. What can be said about oxidizing agent?

A. It loses electron
B. It is oxidized
C. Its charge increases
D. It is reduced

Solution:

Oxidizing agent:

- Causes oxidation
- Gains electrons
- is reduced
- Decreases charge.

Answer: D

39. Given the balanced equation representing an oxidation - reduction reaction:

$$2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$$

Which statement is true about this reaction ?

A. Each Al^{3+} gains $2e^-$ and each Cu loses $3e^-$
B. Each Al^{3+} gains $3e^-$ and each Cu loses $2e^-$
C. Each Al loses $2e^-$ and each Cu^{2+} gains $3e^-$
D. Each Al loses $3e^-$ and each Cu^{2+} gains $2e^-$

Solution:

$Al \rightarrow Al^{3+} + 3e^-$, each Al has lost $3e^-$

$Cu^{2+} + 2e^- \rightarrow Cu$, each Cu^{2+} has gained $2e^-$

Answer: D

40. In a redox reaction :

A. The oxidizing agent undergoes oxidation
B. Oxidation is decrease in oxidation number
C. The oxidizing agent gains electrons.
D. The reduction is increase in oxidation number.

Answer: C

41. Classify the following as oxidation or reduction half reaction.

a. $Cr^{3+} + 3e^- \rightarrow Cr$
b. $N + 3e^- \rightarrow N^{3-}$
c. $2Cl^- \rightarrow Cl_2 + 2e^-$
d. $Sn^{2+} \rightarrow Sn^{4+} + 2e$

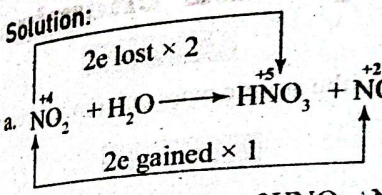
Solution:

- a and b are reduction - half reactions.
- c and d are oxidation - half reactions.

42. Balance the following by oxidation number method.

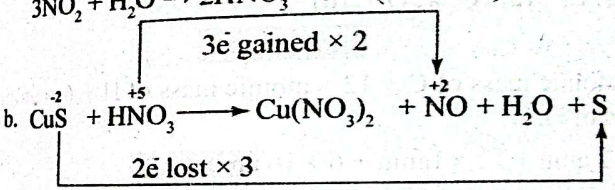
a. $NO_2 + H_2O \rightarrow HNO_3 + NO$
b. $CuS + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O + S$

Solution:



(2+1) $NO_2 + H_2O \rightarrow 2HNO_3 + NO$

$3NO_2 + H_2O \rightarrow 2HNO_3 + NO$ (balanced)



Sometimes we may change the coefficients we have already written.

$$3CuS + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O + 3S$$

43. Identify the reducing and oxidizing agents in the following redox reactions.

a. $2KClO_3 \rightarrow 2KCl + 3O_2$
b. $3MnO_2 + 2Al \rightarrow 3Mn + 2Al_2O_3$

Solution:

	Reducing agent	Oxidizing agent
a.	$KClO_3$	$KClO_3$
b.	Al	MnO_2

44. Write five examples of non-redox reactions.

Solution:

$ZnCO_3 \rightarrow ZnO + CO_2$

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$2NaOH + H_2SO_4 \rightarrow Na_2CO_3 + 2H_2O$

$Ag^+ + Cl^- \rightarrow AgCl$

$K_2SO_4 + BaCl_2 \rightarrow 2KCl + BaSO_4$

$2H_2O \rightarrow H_3O^+ + OH^-$

6.5 Molecular and Formula Masses, the mole concept and chemical formulas

1. Molecular Mass (MM) is the sum of the atomic mass unit (amu) present in a molecule.

Example:

Calculate the molecular mass of glucose ($C_6H_{12}O_6$).

(Atomic mass: C = 12, H = 1, O = 16)

Solution:

MM = 6 × atomic mass of C + 12 × atomic mass of H + 6 × atomic mass of O.

= 6 × 12 amu + 12 × 1amu + 6 × 16 amu

= 180 amu

Atomic mass unit (amu) is a mass exactly equal to one - twelveth the mass of one carbon - 12 atom.

$$1amu = \frac{\text{mass of one C -12 atom}}{12}$$

2. Formula mass (FM) is the sum of the atomic masses of atoms in amu present in a compound (Mainly ionic compound).

Example:

Determine the formula mass of calcium - chloride ($CaCl_2$). (Atomic mass : Cl = 35.5, Ca = 40)

Solution:

FM ($CaCl_2$) = 1 × 40 amu + 2 × 35.5 amu

= 40 amu + 71 amu

= 111 amu

3. The Mole Concept

Mole is the amount of substance that contains as many elementary entities (atoms, molecules, formula units) as there are atoms in exactly 12 grams of C - 12 isotope.

1 mole = 6.02 × 10^{23} Particles

- 6.02×10^{23} is called **Avogadro's number**, in honour of the Italian scientist Amedeo Avogadro.
- Just as one dozen pens contain 12 pens, one mole of hydrogen atoms contains 6.02×10^{23} hydrogen atoms.

12g of C-12 = 1 mol C- 12 atoms

= 6.02×10^{23} C- 12 atoms

Thus, the mass of one C - 12 atom is

$$= 1C-12atom \times \frac{1molC-12atoms}{6.02\times10^{23}C-12atom} \times \frac{12gC-12}{1molC-12atom}$$

$$= 1.993\times10^{-23}\,g$$

Again, it is possible to find the the relationship between atomic mass unit and gram.

$$\frac{gram}{amu} = \frac{1.993\times10^{-23}}{1C-12atom} \times \frac{1Carbon-12atom}{12amu}$$

$$\frac{gram}{amu} = 1.66\times10^{-24}\,g/amu$$

$$\boxed{\begin{aligned} 1amu &= 1.66\times10^{-24}\,g \\ 1g &= 6.02\times10^{23}amu \end{aligned}}$$

Example:

Convert the following

a. 0.5mol of Na into gram

b. 46g of Na into mole (Na = 23g)

Solution:

a) 1mol Na = 23g Na

0.5mol = ?

$$\Rightarrow 0.5\text{mol Na} = \frac{0.5\text{mol Na} \times 23\text{g Na}}{1\text{mol Na}} = 11.5\text{g Na}$$

b) 1mol Na = 23g Na

? = 46g Na

$$\Rightarrow 46\text{g Na} = \frac{1\text{mol Na} \times 23\text{g Na}}{23\text{g Na}} = 2\text{mol Na}$$

4. Molar mass (M) is the mass in grams or kilo grams of one mole of a substance.

For example, the molar mass of C - 12 is 12 g/mol.

Numerically , molar mass of substance is equal to atomic mass of the substance .

$$\text{Number of mole} = \frac{\text{given mass}}{\text{molar mass}}$$

$$\boxed{n = \frac{m}{M}} \quad \boxed{m = n \times M} \quad \boxed{M = \frac{m}{n}}$$

Example:

Calculate the number of moles of 60g NaOH.

(Atomic mass: H = 1, O = 16, Na)

Solution:

$$M = 1\times23 + 1\times16 + 1\times1 = 40\,g/mol$$

$$m = 60g$$

$$n = \frac{m}{M} = \frac{60\,g}{40\,g/mol} = 1.5\,mol$$

Illustrative Examples

45. Calculate the molecular mass of :

a. $C_{12}H_{22}O_{11}$ b. H_3PO_4

(At mass : H = 1, C = 12, O = 16, P = 31)

Solution:

a) MM = 12×12 amu + 22×1 amu + 11×16 amu = 342 amu

b) MM = 3×1amu + 1×31amu + 4×16 amu = 98 amu

46. Find the formula mass of:

a. $Ca_3(PO_4)_2$ b. $Fe_2(SO_4)_3$

(At mass: Ca = 40, Fe = 56, S = 32)

Solution:

a. $Ca_3(PO_4)_2$ here 2 multiplies subscripts of P and O.

FM = 3×40 amu + 2×31 amu + 8×16 amu = 310 amu

b. FM = 2×56 amu + 3×32 amu + 12×16 amu = 400 amu

47. Determine the molar mass of :

a. HNO_3 (At mass : N = 4) b. $MgCO_3$ (At mass : Mg = 24)

Solution:

a. M = 1×1 + 1×14 + 3×16 = 63 g/mol

b. M = 1×24 + 1×12 + 3×16 = 84 g/mol

48. Describe 1 mol of KCl.

Solution:

1 mol KCl = 74.5 g KCl = 6.02×10^{23} formal units of KCl

49. Calculate the number of moles in 1 Kg of calcium carbonate ($CaCO_3$)

Solution:

m = 1kg = 1000g

M = 1×40 + 1×12 + 3×16 = 100g/mol

$$n = \frac{m}{M} = \frac{1000g}{100g/mol} = 10mol$$

50. Find the number of moles of each atom in 2.5 mol of $Al_2(SO_4)_3$

Solution:

Number of mole of atoms in a compound = Mole of compound × Subscript of atom in a compound

= n(Al) = 2 × 2.5mol = 5mol

n(S) = 3 × 2.5mol = 7.5mol

n(O) = 12 × 2.5mol = 30mol

6.6 Chemical, Empirical and Molecular Formulas

1. **Chemical Formula:** is an expression showing the chemical composition of a compound in terms of the symbols for the elements involved.

   Example:

   The chemical formula of water is H_2O. It is composed of two atoms of hydrogen and one atom of oxygen.

2. **Empirical Formula(EF):** is an expression using chemical symbols to show the types of elements in a substance and show the ratios of the different kinds of atoms.

   ➥ It is the simplest formula

3. **Molecular Formula(MF):** is an expression showing the exact number of atoms of each element in molecule.

Table : 6.1 Empirical and Molecular formulas of Some Compounds.

Substance	Molecular formula	Empirical formula
Glucose	$C_6H_{12}O_6$	CH_2O
Benzene	C_6H_6	CH
Hydrogen peroxide	H_2O_2	HO
Butane	C_4H_{10}	C_2H_5
Water	H_2O	H_2O

Percentage composition by mass (%)

$$\%\ \text{Element} = \frac{\text{subscript of element} \times \text{molar mass of element}}{\text{Molar mass of compound}} \times 100$$

Example:

Determine the percent composition by mass of each element in $CaCO_3$.

Solution:

mass of Ca = 1 × 40g = 40g

mass of C = 1 × 12g = 12g

mass of O = 3 × 16g = 48g

mass of $CaCO_3$ = 40 + 12 + 48 = 100g

$$\%\ Ca = \frac{\text{mass Ca}}{\text{mass } CaCO_3} \times 100$$

$$= \frac{40g}{100g} \times 100 = 40\%$$

$$\%\ C = \frac{\text{mass C}}{\text{mass } CaCO_3} \times 100$$

$$= \frac{12g}{100g} \times 100 = 12\%$$

$$\%\ O = \frac{\text{mass O}}{\text{mass } CaCO_3} \times 100$$

$$= \frac{48g}{100g} \times 100 = 48\%$$

Or

% O = 100 - (% Ca + % C)

= 100 - (40 + 12) = 48%

Determination of Empirical Formula

Steps

1. Write down the symbols of the constituent elements
2. Write the mass or mass percentage of the elements under each symbol

3. Find the number of mole of each element
4. To get the simplest whole number values, divide the relative number of atoms, by the smallest of the set.
5. If the above ratio is not whole number, multiply all the values by a suitable integer which will change them to simple whole number.
6. Place the symbols of elements side by side and insert the simplest whole numbers obtained on the lower right corner of the symbols concerned.

Example:

Find the empirical formula of an oxide formed by reaction 28g of iron and 12g of oxygen.

Solution:

Atomic mass: Fe = 56, O = 16

Fe	O
28g	12g
$\frac{28g}{56g/mol}$	$\frac{12g}{16g/mol}$
$\frac{0.5mol}{0.5mol}=1$	$\frac{0.75mol}{0.5mol}$
(1	0.5)×2
2	3

$EF = Fe_2O_3$

Molecular formula = (Empirical formula)x

MF = (EF)x, x is positive integer

Example:

The simplest formula of a compound is CH_2O and its molecular mass is 180. Find the molecular formula of the compound.

Solution:

$MF = (EF)x$

$180 = (CH_2O)x$

$180 = (1\times12 + 2\times1 + 1\times16)x = 30x$

$x = \frac{180}{30} = 6$

$\therefore$ Molecular formula = (Empirical formula)$x = (CH_2O)_6$

$MF = C_6H_{12}O_6$

Illustrative Examples

51. What is the percentage of nitrogen in the fertilizer urea $(CO(NH_2)_2)$

A. 6.7% B. 93.3% C. 46.7% D. 23.3%

Solution:

$CO(NH_2)_2 = CON_2H_4$

$C = 1\times12 = 12g$ $\quad N = 2\times14 = 28g$

$O = 1\times16 = 16g$ $\quad H = 4\times1 = 4g$

Mass $CO(NH_2)_2 = 60g$

$$\%N = \frac{\text{mass N}}{\text{mass } CO(NH_2)_2}\times100 = \frac{28g}{60g}\times100 = 46.7\%$$

Answer: C

52. Which of the following contains the greatest percentage composition of carbon?

A. C_2H_2 B. C_2H_4 C. C_2H_6 D. CH_4

Solution:

$C_2H_2 : 2\times12 + 2\times1 = 26g$

$$\%C = \frac{\text{mass C}}{\text{mass } C_2H_2}\times100 = \frac{24g}{26g}\times100 = 92.3\%$$

$C_2H_4 : 2\times12 + 4\times1 = 28g$

$$\%C = \frac{24g}{28g}\times100 = 85.7\%$$

$C_2H_6 = 2\times12 + 6\times1 = 30g$

$$\%C = \frac{24g}{30g}\times100 = 80\%$$

$CH_4 = 1\times12 + 4\times1 = 16g$

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$$\%C = \frac{12g}{16g} \times 100 = 75\%$$

Answer: A

53. 140 moles of B atoms react with 70 moles of A atoms to give a new compound. The empirical formula of the new compound is _____.

A. B_2A B. BA_2 C. B_2A_3 D. BA

Solution:

B	A
140mol	70mol
$\frac{140}{70} = 2$	$\frac{70}{70} = 1$

$\therefore$ The empirical formula of the compound is B_2A

Answer: A

54. A compound of carbon, hydrogen and oxygen contains 40% C, 6.6%H and 53.4% O. If its molecular formula with a molar mass of 60 is

A. C_2H_2O B. CH_2O C. $C_2H_2O_2$ D. $C_2H_4O_2$

C	H	O
40	6.6	53.4
$\frac{40}{12} = 3.3$	$\frac{6.6}{1} = 6.6$	$\frac{53.4}{16} = 3.3$
$\frac{3.3}{3.3} = 1$	$\frac{6.6}{3.3} = 2$	$\frac{3.3}{3.3} = 1$

$\therefore$ EF = CH_2O

MF = (EF)x

$60 = (CH_2O)x$

$60 = (1\times12 + 2\times1 + 1\times16)x = 30x$

$x = \frac{60}{30} = 2$

$\therefore$ MF = $(CH_2O)_2 = C_2H_4O_2$

Answer: D

55. A compound where analyzed was found to contain 24.58% Potassium, 34.82% manganese and 30.60% oxygen by mass. What is the empirical formula of this compound? (Atomic: K = 39, Mn = 55, O = 16)

Solution:

K	Mn	O
24.5	34.82	40.60
$\frac{24.58}{39} = 0.63$	$\frac{34.82}{55} = 0.63$	$\frac{40.60}{16} = 2.54$
$\frac{0.63}{0.63} = 1$	$\frac{0.63}{0.63} = 1$	$\frac{2.54}{0.63} = 4$

$\therefore$ EF = $KMnO_4$

56. An oxide of nitrogen contains 30.5% nitrogen. Find the empirical of the oxide.

Answer: NO_2

Solve by your own!

57. Calculate the percentage of each element in ammonium nitrate (NH_4NO_3)

Solution:

NH_4NO_3: N: $2\times14 = 28g$

H : $4\times1 = 4g$

O : $3\times16 = 48g$

Mass of NH_4NO_3 = 80g

$$\%N = \frac{\text{mass N}}{\text{mass } NH_4NO_3} \times 100 = \frac{28g}{80g} \times 100 = 35\%$$

$$\%H = \frac{\text{mass H}}{\text{mass } NH_4NO_3} \times 100 = \frac{4g}{80g} \times 100 = 5\%$$

$$\%O = \frac{\text{mass O}}{\text{mass } NH_4NO_3} \times 100 = \frac{48g}{80g} \times 100 = 60\%$$

58. The percentage by mass of water in gypsum is 20.93%. Determine the value of x in the formula of gypsum($CaSO_4 . xH_2O$) (Atomic mass: Ca = 40, S = 32, H = 1, O = 16).

Solution:

$CaSO_4 . x\,H_2O$

$Ca = 1\times40 = 40$

$S = 1\times32 = 32$

$O = 4 \times 16 = 64$

$H_2O = x \times 18 = 18x$

mass of $CaSO_4 . H_2O = 136 + 18x$

$$\%H_2O = \frac{\text{mass } H_2O}{\text{mass } CaSO_4 . H_2O} \times 100$$

$$20.93 = \frac{18x}{136 + 18x} \times 100$$

Solving for x, $x = 2$

∴ Formula of gypsum is $CaSO_4 . 2H_2O$

59. If 18g of carbon is burned in sufficient amount of oxygen to give 66g of carbon dioxide, what is the percentage of oxygen in carbon dioxide?

Solution:

$$C + O_2 \longrightarrow CO_2$$

18g x 66g

$18g + x = 66g$

$x = 66g - 18g = 48g = \text{mass O}$

$$\%\text{ of O} = \frac{\text{mass O}}{\text{mass } CO_2} \times 100 = \frac{48g}{66g} \times 100 = 72.7\%$$

60. Mercuric oxide contains 92.6% mercury by mass Calculate the mass of mercuric oxide which contains 10g of oxygen.

Solution:

$\%\text{of O} = 100 - \%Hg = 100 - 92.6 = 7.4\%$

$$\%O = \frac{\text{mass O}}{\text{mass HgO}} \times 100$$

$$\text{mass HgO} = \frac{\text{mass O}}{\%O} \times 100$$

$$= \frac{10g}{7.4\%} \times 100 = 135g$$

6.7 Stoichiometry

The study of the quantitative relationship between reactant and reactant, reactant and product or product and product is known as **stoichiometry.**

Types of Stoichiometric Relationship

1. Mass - Mass Relationships

- Write the balanced chemical equation
- Write the given mass above the formula and x above the required.
- Below the formula, write coefficient times molar mass.
- Set a proportion and solve for x

Example:

How many grams of ammonia is formed when 7g of nitrogen reacts with excess hydrogen. (m: $N_2 = 28g$, $NH_3 = 17\,g$)

Solution:

$$\overset{7g}{N_2} + 3H_2 \longrightarrow 2\overset{x}{NH_3}$$

$1 \times 28g$ $2 \times 17g$

$$\frac{7g}{1 \times 28g} = \frac{x}{2 \times 17g}$$

$\frac{7g}{28g} = \frac{x}{34g} = 8.5g$ of NH_3 Thus mass of NH_3 produced is 8.5g

2. Mole - mass Relationship

Below mole, write coefficient and below mass coefficient times molar mass

Solution

$$2\overset{x}{CO} + \overset{0.5mol}{O_2} \longrightarrow 2CO_2$$

$2 \times 28g$ $1mol$

$$\frac{x}{2x28g} = \frac{0.5mol}{1} \Rightarrow x = 28g \text{ of CO}$$

Example

Determine the mass of CO that reacts with 0.5mol of O_2.

Thus, the mass of CO that reacts with 0.5mol of O_2 is 28g.

3. Mole - mole Relationship

Given mole is written above the formula and coefficient below the formula as number of mole.

Example:

How many moles of hydrochloric acid reacts with excess aluminum to form 6 moles of hydrogen.

Solution

$$2Al + \underset{6mol}{6HCl} \longrightarrow 2AlCl_3 + \underset{3mol}{3H_2}$$

$$\frac{x}{6mol} = \frac{6mol}{3mol} = 12mol \text{ of } HCl$$

Thus, the number of HCl required is 12mol.

4. Mass - Volume Relationship

Molar volume is the volume of one mole of any gas at standard temperature and pressure (STP)

➥ At STP, temperature = 0°C, pressure = 1 atm, number of mole = 1 mol and volume = 22.4L

➥ Write the volume above the formula and coefficient times 22.4L below the formula.

➥ For mass, follow the same method.

Example:

How many liters of carbon dioxide gas is formed when 50g of calcium carbonate is decomposed at STP? (M: $CaCO_3$=100g)

Solution:

$$\overset{50g}{\underset{1\times100g}{CaCO_3(s)}} \longrightarrow CaO(s) + \overset{x}{\underset{1\times22.4g}{CO_2(g)}}$$

$$\frac{50g}{100g} = \frac{x}{22.4g}, x = 11.2L$$

Thus, the volume of CO_2 formed is 11.2L.

5. Mole - Volume Relationship

Below mole write coefficient and below volume coefficient times 22.4L.

How many liters of oxygen gas reacts with 3moles of aluminum to form aluminium oxide at STP?

Solution:

$$\underset{4mol}{\overset{3mol}{4Al(s)}} + \underset{3\times22.4L}{\overset{x}{3O_2}}(g) \longrightarrow 2Al_2O_3(s)$$

$$\frac{3mol}{4mol} = \frac{x}{3\times22.4L} = \text{solving for } x = 50.4L$$

Thus the volume of oxygen is 50.4L

6. Volume - Volume Relationship

Write coefficient below the formulas of the given and required substances.

Example:

How many litter of carbon dioxide gas will be formed when 44.8 litters of acetylene is burnt in excess oxygen.

Solution:

$$\underset{2}{\overset{44.8L}{2C_2H_2}}(g) + 5O_2(g) \longrightarrow \underset{4}{\overset{x}{4CO_2}} + 2H_2O(l)$$

$$\frac{44.8L}{2} = \frac{x}{4}, x = 89.6L$$

Thus, the volume of CO_2 formed is 89.6L.

Illustrative Examples

61. Given the reaction : $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(g)$. How many grams of Fe_2O_3 are produced when 32g of iron reacts with sufficient amount of oxygen?

A. 12g B. 45.7g C. 91.4g D. 106.7g

Solution:

$$\overset{32g}{\underset{4\times56g}{4Fe}} + 3O_2 \longrightarrow \underset{2\times160g}{\overset{x}{Fe_2O_3}}$$

$$\frac{32g}{4\times56g} = \frac{x}{2\times160g}, x = 45.7g \text{ of } Fe_2O_3$$

Answer: B

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62. Given the reaction : $2KNO_3 \rightarrow 2KNO_2 + O_2$, how many moles of potassium nitrate must be decomposed to produce 9 moles of oxygen gas?

A. 18 moles B. 3 moles C. 9 moles D. 2 moles

Solution:

$$\overset{x}{2KNO_3} \longrightarrow 2KNO_2 + \overset{9mol}{O_2}$$
$$\underset{2mol}{\quad} \qquad \underset{1mol}{\quad}$$

$$\frac{x}{2mol} = \frac{9mol}{1mol}, x = 18mol \text{ of } O_2$$

Answer: A

63. How many liters of carbon monoxide at STP are required to react with 11.2L of hydrogen gas according to the equation below?

$CO(g) + 2H_2(g) \rightarrow CH_3OH(l)$

A. 44.8L B. 22.4L C. 11.2L D. 5.6L

Solution:

$$\overset{x}{CO}(g) + \overset{11.2L}{2H_2}(g) \longrightarrow CH_3OH(l)$$
$$\underset{1}{\quad} \qquad \underset{2}{\quad}$$

$$\frac{x}{1} = \frac{11.2L}{2}, x = 5.6L \text{ of } CO$$

Answer: D

64. The equation for the reaction of sucrose with oxygen as follow:

$C_{12}H_{22}O_{11} + 12O_2 \rightarrow 12CO_2 + 11H_2O$

How many litters of carbon dioxide gas is produced at STP per gram of sucrose?

Solution:

Per gram means per one gram

$$\underset{1\times342g}{\overset{1g}{C_{12}H_{22}O_{11}}} + 12O_2 \longrightarrow \underset{12\times22.4l}{\overset{x}{12CO_2}} + 11H_2O$$

$$\frac{1g}{342g} = \frac{x}{12\times22.4L}, x = 0.786L \text{ of } CO_2$$

Thus, the volume of CO_2 formed is 0.786L

65. Given the reaction:

$CaC_2(s) + 2H_2O(1) \rightarrow C_2H_2(g) + Ca(OH)_2(aq)$,

What mass of CaC_2 is required to produce 5600ml of C_2H_2 at STP? (5600ml = 5.6L)

Solution:

$$\underset{1\times64g}{\overset{x}{CaC_2}}(s) + 2H_2O(1) \longrightarrow \underset{1\times22.4L}{\overset{5.6L}{CH}}(g) + Ca(OH)_2(aq)$$

$$\frac{x}{64g} = \frac{5.6L}{22.4L}, x = 16g \text{ of } CaC_2$$

66. The food we eat is degraded or broken down in our bodies to provide energy for growth and function. A general overall equation for this complex process represents the degradation of glucose to carbon dioxide and water.

$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + H_2O$

If 856g of $C_6H_{12}O_6$ are consumed by the body over a certain period, What is the mass of CO_2 produce?

Solution:

$$\underset{1\times180g}{\overset{856g}{C_6H_{12}O_6}} + \overset{x}{6O_2} \longrightarrow \underset{6\times44g}{\overset{x}{6CO_2}} + 6H_2O$$

$$\frac{856g}{180g} = \frac{x}{6\times44g}, x = 1255.5g = 1.25kg$$

Thus, amount of CO_2 produce was 1.25kg

67. What is the importance of balanced chemical equation in solving stoichiometric calculations?

Solution:

Balanced chemical equation is used to find the mole ratio of reactants and products.

6.8 Limiting Reactant, Excess Reactant, Theoretical yield, Actual and Percentage Yield

☑ When a chemist carries out a reaction, the reactants are not usually present in exact **Stoichiometric amounts**, that is, in the proportions

indicated by the balanced equation.

- ☑ The reactant used up first in a reaction is called the **limiting reactant.**
- ☑ The amount of product formed depends on **limiting reactant.**
- ☑ The other reactant present in quantities greater than those needed to react with the quantity of limiting reactant are called **excess reactant.**

Note:

To identify the limiting reactant, first determine the total number of moles (n_T) of reactants.

The one with smaller n_T is limiting reactant.

$$n_T = \frac{\text{Given mass}}{\text{Coefficient} \times \text{molar mass}}$$

$$n_T = \frac{m}{\text{coeff} \times M}$$

Example:

Given the reaction:

$2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$.

If a mixture of 108g Al and 240g of Fe_2O_3 were ignited.

a. Identify the limiting reactant.

b. Calculate the maximum amount of iron produced.

c. How many grams of excess reactant left at the end of the reaction?

Solution:

a) $n_T(Al) = \frac{m}{\text{coeff.} \times M} = \frac{108g}{2 \times 27g/mol} = 2mol$

$$n_T(Fe_2O_3) = \frac{m}{\text{coeff.} \times M} = \frac{240g}{1 \times 160g/mol} = 1.5mol$$

∴ Therefore, Fe_2O_3 is limiting reagent because it has smaller n_T.

b) The amount of product depends on the amount of limiting reactant.

$$\overset{240g}{2Al} + \underset{1\times160g}{Fe_2O_3} \longrightarrow Al_2O_3 + \underset{2\times56g}{\overset{x}{2Fe}}$$

$$\frac{240g}{160g} = \frac{x}{2 \times 56g}, \quad x = 168g \text{ of Fe}$$

∴ Thus, 168g of Fe is formed.

c) Al is the excess reactant. Find the mass of Al that reacts with 240g and then subtract this value from the given mass of Al.

$$\overset{x}{2Al} + \underset{1\times160g}{\overset{240g}{FeO_3}} \longrightarrow Al_2O_3 + 2Fe$$

$$\frac{x}{2\times27g} = \frac{240g}{1\times160g}, \quad x = 81g \text{ of Al is reacted}$$

∴ Mass of Al left = 108g - 81g = 27g

Percentage yield, Theoretical yield and Actual yield

$$\text{Percentage yield} = \frac{\text{Actual yiled}}{\text{Theoretical yiled}} \times 100$$

$$\%\text{ yield} = \frac{AY}{TY} \times 100$$

$$TY = \frac{AY}{\%\text{yiled}} \times 100$$

$$AY = \frac{\%\text{yield} \times TY}{100}$$

AY - Actual yield, TY & Theoretical yield

- **Actual yield** is the amount of product determined experimentally
- **Theoretical yield** is the amount of product obtained stoichiometrically from balanced chemical equation.

Actual yield is always less than or equal to theoretical yield.

Example:

A grade nine student heated 99.8g of copper in excess sulphur and obtained 80g of copper (I) sulphide (Cu_2S). Determine the percentage yield.

$$2Cu(s) + S(s) \rightarrow Cu_2S(s)$$

Solution:

Actual yield = 80g

Calculate the theoretical yield from the balanced chemical equation.

$$\underset{2\times64g}{\overset{99.8g}{2Cu}} + S \longrightarrow \underset{1\times160g}{\overset{TY}{Cu_2S}}$$

$$\frac{99.8g}{2\times64g} = \frac{TY}{1\times160g} \Rightarrow Ty = 124.75g$$

$$\%yield = \frac{AY}{TY}\times100 = \frac{80g}{124.75g}\times100 = 64.1\%$$

Illustrative Examples

68. Which of the following is true about the theoretical yield, actual and percentage yield?

A. Percentage yield = Theoretical yield ×100

B. When the amount of product is calculated from a balanced chemical equation the result is known as actual yield

C. The actual yield is less than the theoretical yield

D. The theoretical yield is the minimum amount of the product that can be produced from the quantities of the reactant used

Solution:

- % Yield = $\frac{AY}{TY} \times 100$
- Actual yield is practically determined in the laboratory
- Theoretical yield is calculated from balanced chemical equation
- $AY \leq TY$

Answer: C

69. When 32.5g of Zinc reacted with excess hydrochloric acid, 0.4mol of hydrogen gas was formed, what was the percentage yield of hydrogen?

A. 80% B. 70% C. 60% D. 40%

Solution:

$$\underset{1\times65g}{\overset{32.5g}{Zn}} + 2HCl \longrightarrow ZnCl_2 + \underset{1\times2g}{\overset{x}{H_2}}$$

$$\frac{32.5g}{65g} = \frac{x}{2g} \Rightarrow x = TY = 1g \text{ of } H_2$$

$$AY = 0.4mol = 0.4mol \times \frac{2g}{mol} = 0.8g$$

$$\%yield = \frac{AY}{TY}\times100 = \frac{0.8g}{1g}\times100 = 80\%$$

Answer: A

70. Given the balanced equation that represents the reaction: $NaOH + HCl \rightarrow NaCl + H_2O$. What mass of sodium chloride is formed when 60g of NaOH is mixed with 73g of HCl?

A. 39g B. 36.5g C. 117g D. 87.8g

Solution:

First identify the limiting reactant because the mass of product formed depends on it.

$$n_T(NaOH) = \frac{m}{Coeff\times M} = \frac{60g}{1\times40g/mol} = 1.5mol$$

$$n_T(HCl) = \frac{m}{Coeff\times M} = \frac{73g}{1\times36.5g/mol} = 2mol$$

$\Rightarrow$ NaOH is limiting reactant

$$\underset{1\times40g}{\overset{60g}{NaOH}} + HCl \longrightarrow \underset{1\times58.5g}{\overset{x}{NaCl}} + H_2O$$

$$\frac{60g}{40g} = \frac{x}{58.5g} \Rightarrow x = \text{mass of NaCl} = 87.75g$$

Answer: D

71. 2 moles of magnesium is heated in 3 moles of oxygen. How many moles of the excess reactant are left unreacted according to the following equation? $2Mg + O_2 \rightarrow 2MgO$

A. 1mol O_2 B. 1.5mol Mg C. 2mol O_2 D. 1mol Mg

Solution:

$$n_T(Mg) = \frac{2mol}{2} = 1mol$$

$$n_T(O_2) = \frac{3mol}{1} = 3mol$$

$\Rightarrow$ Mg is the limiting reactant.

The number of moles of O_2 reacted is 1mole.

The unreacted number of moles is 3 - 1 = 2 moles **Answer: C**

72. 1.6g of CaC_2 and 1.8g of H_2O were mixed and made to react according to the following reaction, $CaC_2 + 2H_2O \rightarrow C_2H_2 + Ca(OH)_2$

The limiting reactant is:

A. CaC_2 B. H_2O C. C_2H_2 D. $Ca(OH)_2$

Solution:

$$n_T(CaC_2) = \frac{1.6g}{1\times 64g/mol} = 0.025mol$$

$$n_T(H_2O) = \frac{1.8g}{2\times 18g} = 0.05mol$$

$\therefore$ CaC_2 is the limiting reactant.

Answer: A

73. A 25g sample of aluminum sulphide contains 8.98g of aluminium. Calculate the percentage of sulphur in the sample.

Solution:

mass of sulphur = mass of sample - mass of aluminum

= 25g - 8.98g = 16.02g

$$\%S = \frac{\text{mass of S}}{\text{mass of sample}} \times 100$$

$$= \frac{16.02g}{25g} \times 100 = 64\%$$

74. If 5g of sulphur dioxide is actually formed from the reaction of 3g oxygen with sufficient sulphur, what is the percentage yield of SO_2?

Solution:

$$\underset{1\times 32g}{\overset{3g}{S}} + O_2 \longrightarrow \underset{1\times 64g}{\overset{TY}{SO_2}}$$

$$\frac{3g}{32g} = \frac{TY}{64g} \Rightarrow TY = 6g$$

$$\%yield = \frac{AY}{TY} \times 100$$

$$= \frac{5g}{6g} \times 100 = 83.3\%$$

75. When haematite is reduced to iron is a blast furnace, the percentage yield of iron is 78%. Calculate the actual mass of iron in kilogram that would be formed from the reduction of $1\times 10^9 g$

Solution:

$$\underset{1\times 160g}{\overset{1\times 10^9 g}{Fe_2O_3}} + 3CO \longrightarrow \underset{2\times 56g}{\overset{TY}{2Fe}} + 3CO_2$$

$$\frac{1\times 10^9 g}{160g} = \frac{TY}{2\times 56g} \Rightarrow TY = 7\times 10^8 g$$

$$AY = \frac{\%yield \times TY}{100} = \frac{78\times 7\times 10^8 g}{100} = 5.46\times 10^8 \text{ grams}$$

The actual mass of iron produced is 5.46×10^5 kg.

Solved Examples

76. List some indications for the occurrence of chemical change.

Solution:

- Formation of new substances.
- Production of heat or light or both.
- Change in colour.
- Change in temperature

77. Write the chemical equation for the reaction between calcium hydroxide and nitric acid.

Solution:

Calcium hydroxide + nitric acid → Calcium nitrate + water

$Ca(OH)_2(aq) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2H_2O(l)$

78. Explain the qualitative and quantitative meaning of the balanced chemical equation below. $N_2 + 3H_2 \rightarrow 2NH_3$

Solution:

Qualitatively:- it indicates the reaction between nitrogen and hydrogen that results ammonia.

Quantitatively:- it show that

- One unit of nitrogen reacts with three units of hydrogen to form two units of ammonia.
- 28g of nitrogen reacts with 6g of hydrogen to give 34g of ammonia.
- 1 mole of nitrogen reacts with 3moles of hydrogen to pro-

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duce 2 moles of ammonia.

➥ 6.02×10^{23} molecules of nitrogen reacts with 1.806×10^{24} molecules of hydrogen to form 1.204×10^{24} molecules of ammonia.

79. Balance the following equations using algebraic method.

a. $P_4O_6 + H_2O \rightarrow H_3PO_3$ b. $Al + HBr \rightarrow AlBr_3 + H_2$

Solution:

a. $a\,P_4O_6 + b\,H_2O \rightarrow c\,H_3PO_3$ assume a = 1

P: $4a = c$ $\Rightarrow c = 4\times1 = 4$

O: $6a + b = 3c$ $2b = 3c$

H: $2b = 3c$ $2b = 3\times4$

$\Rightarrow b = 6$

$\therefore P_4O_6 + 6H_2O \rightarrow 4H_3PO_3$

b. $a\,Al + b\,HBr \rightarrow c\,AlBr_3 + d\,H_2$ Br: $b = 3c$

Al: $a = c$ Br: $b = 3c$

Assume $a = 1$ $b = 3c = 3\times1 = 3$

H: $b = 2d$ $b = 2d = 3 = d = 3/2$

$$\left.\begin{matrix} a=1 \\ b=3 \\ c=1 \\ d=3/2 \end{matrix}\right] \times 2 \Rightarrow \begin{matrix} a=2\times1=2 \\ b=3\times2=6 \\ c=1\times2=2 \\ d=3/2\times2=3 \end{matrix}$$

$\therefore 2Al + 6\,HBr \rightarrow 2AlBr_3 + 3H_2$

80. Given the reaction: $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$ Identify

a. Reduced element c. Oxidizing agent

b. Oxidized element d. Reducing agent

Solution:

$$\overset{+3}{Fe_2O_3} + 3\overset{+2}{C}O \longrightarrow 2\overset{0}{Fe} + 3\overset{+4}{C}O_2$$

a. Fe because its oxidation number is decreased.

b. C because its oxidation number is increased.

c. Fe_2O_3 Since it contains reduced element.

d. CO because it has Oxidized element.

81. Find the oxidation number of

a. Br in $HBrO_3$ c. N in NO_3^-

b. S in $Fe_2(SO_4)_3$ d. P in PH_4^+

Solution:

a. $\overset{+1}{H}\overset{x}{Br}\overset{-2}{O_3}$

$x + 1 + 3(-2) = 0$

$x = +5$

$\therefore$ The oxidation number of Br in $HBrO_3$ is +5

b. $\overset{+3}{Fe_2}\left(\overset{x}{SO_4}^{-2}\right)_3$

$2(+3) + 3x + 12(-2) = 0 \Rightarrow x = +6$

$\therefore$ The oxidation state of S in $Fe_2(SO_4)_3$ is +6.

c. $\overset{x}{NO_3^-}$

$x + 3(-2) = -1,\ x = +5$

$\therefore$ The oxidation state of N in NO_3^- is +5

d. $\overset{x}{PH_4^+}$

$x + 4(+1) = +1,\ x = -3$

$\therefore$ The oxidation state of P in PH_4^+ is -3

82. Explain the colour changes that take place in acidic solution of the following strong oxidizing agent ions.

a. MnO_4^- in acidic solution changes colour from purple to colourless.

$$\underset{\text{purple}}{2MnO_4^-} + 16H^+ + 10e^- \longrightarrow \underset{\text{Colourless}}{2Mn^{2+}} + 8H_2O$$

b. $Cr_2O_7^{2-}$ in acidic solution changes colour from orange to green.

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow \overset{3+}{2Cr} + 7H_2O$$

organge green

83. Balance the following redox reactions using oxidation number method.

a. $K_2Cr_2O_7 + FeCl_2 + HCl \longrightarrow CrCl_3 + FeCl_3 + KCl + H_2O$

b. $MnO_2 + Al \longrightarrow Mn + Al_2O_3$

Solution:

First balance the reduced and oxidized elements

a. $K_2Cr_2O_7 + FeCl_2 + HCl \longrightarrow 2CrCl_3 + FeCl_3 + KCl + H_2O$

1e⁻ loss ×6

6e⁻ gained ×1

Balance the rest by inspection method.

$\Rightarrow K_2Cr_2O_7 + 6FeCl_2 + 14HCl \longrightarrow$

(balanced) $2CrCl_3 + 6FeCl_3 + 2KCl + 7H_2O$

b. $\overset{+4}{MnO_2} + 2\overset{0}{Al} \longrightarrow \overset{0}{Mn} + \overset{+3}{Al_2O_3}$

6e⁻ loss ×2

4e⁻ gained ×3

$\Rightarrow 3MnO_2 + 4Al \longrightarrow 3Mn + 2Al_2O_3$

(balanced)

84. Determine the number of atoms and ions in one formula unit of

a. Na_2CO_3 b. $Ca_3(PO_4)_2$

Solution:

In one formula unit:

➥ To determine the number of ions, write the dissociation or ionization reaction.

➥ To find number of atoms, simply add the subscripts.

a. $Na_2CO_3 \longrightarrow 2\overset{+}{Na} + \overset{2-}{CO_3}$

Number of ions $= 2(Na^+) + 1(CO_3^{2-}) = 3$

Number of atoms $= 2Na + 1C + 3O = 6$

b. $Ca_3(PO_4)_2 \longrightarrow 3\overset{2+}{Ca} + 2\overset{3-}{PO_4}$

Number of ions $= 3(\overset{2+}{Ca}) + 2(\overset{3-}{PO_4}) = 5$

Number of atoms $= 3(Ca) + 2(P) + 8(O) = 13$

85. Determine the number of mole each element in 2mol $Fe_3(PO_4)_2$.

Solution:

Number of mole of an element in a compound = number of mole of compound × Subscript of element

$\Rightarrow n(Fe) = 2mol \times 3 = 6mol$

$n(P) = 2mol \times 2 = 4mol$

$n(O) = 2mol \times 8 = 16mol$

86. Arrange the following in increasing order of number of molecules.

(i) 1g of CH_4 (ii) 1g of O_2 (iii) 1g of H_2 (iv) 1g of H_2O

Solution:

$$\text{Number of molecules} = \frac{m}{M} \times 6.02 \times 10^{23}$$

i. Number of molecules $(CH_4) = \frac{1g}{16g/mol} \times 6.02 \times 10^{23}$

$$= 3.76 \times 10^{22} \frac{\text{molecules}}{\text{mol}}$$

ii. Number of molecules $(O_2) = \frac{1g}{32g/mol} \times 6.02 \times 10^{23}$

$$= 1.88 \times 10^{22} \frac{\text{molecules}}{\text{mol}}$$

iii. Number of molecules $(H_2) = \frac{1g}{2g/mol} \times 6.02 \times 10^{23}$

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$$= 3.01\times10^{23}\ \frac{\text{molecules}}{\text{mol}}$$

iv. Number of molecules $(H_2O) = \frac{1g}{18g/mol}\times 6.02\times10^{23}$

$$= 3.34\times10^{22}\ \frac{\text{molecules}}{\text{mol}}$$

$\therefore$ ii < iv < i < iii

87. A sample of glucose, $C_6H_{12}O_6$ contains 2×10^{20} atoms of oxygen. How many atoms of hydrogen does it contain?

Solution:

Use subscript ratio method to determine the answer.

H : O

12 : 6

2 : 1

? : 2×10^{20}

$\Rightarrow$ No of H - atoms $= \frac{2\times2\times10^{20}}{1} = 4\times10^{20}$ atoms

88. When 6.5g of zinc were added to solution of copper (II) sulphate, 6.4g of copper were obtained.

a. Write balanced chemical equation
b. What type of chemical reaction is it?
c. How many zinc atoms were used?
d. How many grams of copper (II) sulphate were consumed?
e. How many grams of zinc sulphate were formed?

Solution:

a. $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

b. It is single displacement reaction

c. Number of Zn atoms $= \frac{m(Zn)}{M(Zn)}\times 6.02\times10^{23}$

$$= \frac{6.5g}{65g}\times 6.02\times10^{23} = 6.02\times10^{22}\text{ atoms}$$

d. $\underset{1\times65g}{\overset{6.5g}{Zn}} + \underset{1\times160g}{\overset{x}{CuSO_4}} \longrightarrow ZnSO_4 + Cu$

$$\frac{6.5g}{65g} = \frac{x}{160g}$$

x = masss of $CuSO_4$ consumed = 16g

e. $\underset{1\times65g}{\overset{6.5g}{Zn}} + CuSO_4 \longrightarrow \underset{1\times161g}{\overset{x}{ZnSO_4}} + Cu$

$$\frac{6.5g}{65g} = \frac{x}{161g}$$

$x = 16.1g$ of $ZnSO_4$ Were formed.

89. A 1.871 g sample of unknown metallic carbonate is decomposed by heating to form the metallic oxide and 0.656g of carbon dioxide according to the equation: $MCO_3 \rightarrow MO + CO_2$

Identify the metal M.

Solution:

To identify the metal, we have to find the atomic mass of the metal based on the given data.

$$\underset{1.871g}{MCO_3} \longrightarrow \underset{x}{MO} + \underset{0.656g}{CO2}$$

$1.871g = x + 0.656g$

x = mass of MO = 1.871-0656 = 1.215g

By using mass - mass relationship

$$\frac{\text{mass MO}}{\text{molar mass MO}} = \frac{\text{mass } CO_2}{\text{molar mass } CO_2}$$

Let the atomic mass of M be x

$$\Rightarrow \frac{1.215g}{(x+16)g} = \frac{0.656g}{1\times44g}, \text{solving for } x,\ x = 65g/mol$$

Thus, the metal is Zinc.

90. What are the number of moles and mass of calcium in a sample containing 1.204×10^{24} atoms?

Solution:

1mole Ca = 6.02×10^{23} atoms

x ? $= 1.204\times10^{23}$ atoms

$$x = n = \frac{1mol \times 1.204\times10^{24}\ atoms}{6.02\times10^{23}\ atoms} = 2\ moles$$

$$\Rightarrow m = n \times M$$

$$m = 2mol \times \frac{40g}{mol} = 80g$$

91. What mass of magnesium contains the same number of atoms as 5g of calcium?

Solution:

$40g\ Ca = 6.02\times10^{23}$ atoms

$5g\ Ca = ?\ x$

$$x = \frac{5g\ Ca \times 6.02\times10^{23}\ atoms}{40g\ Ca}$$

$x = 7.525\times10^{23}$ Ca atoms.

$$\text{Number of Ca - atoms} = \text{Number of Mg - atoms} = 7.525\times10^{23}\ \text{atoms.}$$

$\Rightarrow 24g = 6.02\times10^{23}$ atoms.

x ? $= 7.525\times10^{22}$ atoms

$$x = \frac{24g \times 7.525\times10^{22}\ atoms.}{6.02\times10^{23}\ atoms.} = 3.01g$$

Thus, the mass of Mg is 3.01g

92. 60g of a metal M (M = 60) combine with 24g of oxygen (M = 16) to form an oxide. What is the formula of the oxide?

Solution:

M	O
$\frac{60}{60}$	$\frac{24}{16}$
(1	1.5)×2
2	3

∴ The formula of oxide is M_2O_3

93. What is the mass of one molecule of ethane (C_2H_6)?

Solution:

Molecular mass of $C_2H_6 = 6.02\times10^{23}$ molecules

$\Rightarrow 30g\ C_2H_6 = 6.02\times10^{23}$ molecule

x ? = one molecule

$$x = \frac{30gC_2H_6 \times molecule}{6.02\times10^{23}\ molecules}$$

$x = 4.98\times10^{-23}g$

Thus, the mass of one molecule of C_2H_6 is 4.98×10^{-23}g.

94. What is the number of chlorine atoms in 1.0×10^{-5}mol of CCl_4?

Solution:

$$\text{Number of atoms of an element in a compound} = n(\text{Compound}) \times \text{Subscript of element} \times 6.02\times10^{23}$$

$$\text{No of Cl-atoms} = nCCl_4 \times 4 \times 6.02\times10^{23}$$

$$= 1.0\times10^{-5}\ mol \times 4 \times 6.02\times10^{23}\ \frac{atoms}{mol}$$

$$= 2.408\times10^{19}\ atoms$$

95. How many Chloride ions are there in 2moles of $MgCl_2$?

Solution:

$$\text{Number of ions in ionic compound} = n(\text{Compound}) \times \text{Subscript of ion} \times 6.02\times10^{23}$$

$$\text{No of } Cl^- = 2mol \times 2 \times 6.02\times10^{23}\ \frac{ions}{mol}$$

$$= 2.408\times10^{24}\ ions$$

96. How many atoms are present in 64g of SO_3?

Solution:

$$\text{Number of atoms of a compound} = \frac{m}{M} \times \text{Sum of Subscript} \times 6.02\times10^{23}\ \frac{atoms}{mol}$$

$$\text{Number of atoms in 64g of } SO_3 = \frac{64g}{80g/mol} \times (1+3) \times 6.02 \times 10^{23} \frac{\text{atoms}}{\text{mol}}$$

$$= 1.93 \times 10^{24} \text{ atoms}$$

97. 3moles of $(NH_4)_2HPO_4$ contains_____ moles of hydrogen atoms.

$$\text{Number of moles of atom in a compound} = \text{n(Compound)} \times \text{Subscript of element}$$

$$\Rightarrow n(H) = 3mol \times 9 = 27mol$$

98. How many moles of water will be produced from complete combustion of 2moles of benzene, C_6H_6?

Solution:

First, we have to write the correct balanced chemical equation and find the answer by mole-mole relationship.

$$\overset{2mol}{2}C_6H_6 + 15O_2 \longrightarrow \overset{x}{6}H_2O + 12CO_2$$

$$\frac{2mol}{2} = \frac{x}{6} \Rightarrow = 6mol$$

Thus, 6 moles of water will be produced.

99. How many grams of CO_2 would be formed from the complete combustion of 3moles of ethane, C_2H_6?

Solution:

$$\overset{3mol}{2}C_6H_6 + 7O_2 \longrightarrow \overset{x}{4}CO_2 + 6H_2O$$

$$\frac{3mol}{2mol} = \frac{x}{4 \times 44g} \quad x = 264g \text{ of } CO_2$$

Thus, the mass of CO_2 formed is 264g.

100. A chemist wants to make 20kg of methane, CH_4. How many molecules is that?

Solution:

Molar mass of $CH_4 = 6.02 \times 10^{23}$ molecules

$\Rightarrow 16g$ of $CH_4 = 6.02 \times 10^{23}$ molecules

$20kg = 20,000g = ?x$

$$x = \frac{20,000g \times 6.02 \times 10^{23} \text{ molecules}}{16g} = 7.525 \times 10^{26} \text{ molecules}$$

$\therefore 20kg = 7.525 \times 10^{26}$ molecules

101. What is the mass of SO_3 produced when 48g of SO_2 and 12g of O_2 are reacted?

Solution:

$$2SO_2 + O_2 \longrightarrow 2SO_3$$

Identify the limiting reactant.

$$n_T(SO_2) = \frac{m}{Coeff.M} = \frac{48g}{2 \times 64g/mol} = 0.375mol$$

$$n_T(O_2) = \frac{m}{Coeff.M} = \frac{12g}{1 \times 32g/mol} = 0.375mol$$

$\therefore$ Both SO_2 and O_2 are limiting reactants.

$$\overset{48g}{2SO_2} + O_2 \longrightarrow \overset{x}{2SO_3}$$
$$\underset{2\times64g}{} \qquad \underset{2\times80g}{}$$

$$\frac{48g}{2 \times 64g} = \frac{x}{2 \times 80g} \quad x = 60g$$

$\therefore$ Therefore mass of SO_3 formed is 60g.

102. Give the reaction: $3NaBH_4 + 4BF_3 \rightarrow 3NaBF_4 + 2B_2H_6$ If the reaction has a percentage yield of 70%, How many moles of NaB_4 should be used with excess BF_3 in order to obtain 0.2mol of B_2H_6?

Solution:

$\%yield = 70\%$

$AY = 0.2mol$

$$TY = \frac{AY}{\%yield} \times 100$$

$$= \frac{0.2mol}{70} \times 100 = 0.286mol \text{ of } B_2H_6$$

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By mole - mole relationship, find the number of mole of $NaBH_4$

$$3NaBH_4 + 4BF_4 \longrightarrow \overset{x}{3NaBH_4} + \overset{0.286mol}{2B_2H_6}$$
$$\qquad\qquad\quad 3mol \qquad\quad 2mol$$

$$\frac{x}{3mol} = \frac{0.286mol}{2mol},\ x = 0.429mol \text{ of } NaBF_4 \text{ is formed.}$$

103. How many grams of solid KO_2 is needed to react with 50L of CO_2 at STP according to the reaction below:

$$4KO_2(s) + 2CO_2(g) \rightarrow 2K_2CO_3(s) + 3O_2(g)$$

Solution:

$$\overset{x}{4KO_2}(s) + \overset{50L}{2CO_2}(g) \longrightarrow 2K_2CO_3(s) + 3O_2(g)$$
$$\qquad 4\times71g \qquad\quad 2\times22.4L$$

$$\frac{x}{4\times71g} = \frac{50L}{2\times22.4L},\ x = 316.96g \text{ of } KO_2$$

104. Ten grams of Calcium carbonate was decomposed to give 3.96g of Carbon dioxide according to an equation representing the reaction.

$$CaCO_3 \rightarrow CaO + CO_2$$

Solve by your own! **Answer: 90%**

105. How many grams of $Fe(OH)_3$ can be produced by reacting 1.0 mole Fe_2S_3, 2 moles of H_2O and 3 moles of O_2, according to the following Chemical equation?

$$2Fe_2S_3 + 6H_2O + 3O_2 \rightarrow 4Fe(OH)_3 + 6S$$

Solution:

Divide the number of mole by the coefficient to know the limiting reactant.

$$\Rightarrow n(Fe_2S_3) = \frac{1mol}{2} = 0.5mol$$

$$n(H_2O) = \frac{2mol}{6} = 0.33mol$$

$$n(O_2) = \frac{3mol}{3} = 1mol$$

$\Rightarrow H_2O$ is limiting reactant

$$2Fe_2S_3 + \overset{2mol}{6H_2O} + 3O_2 \longrightarrow \overset{x}{4Fe(OH)_3} + 6S$$
$$\qquad\qquad 6mol \qquad\qquad 4\times107g$$

$$\frac{2mol}{6mol} = \frac{x}{4\times107g} = x = 142.7g \text{ of } Fe(OH)_3$$

106. Calculate the molar mass of a compound if 0.25mol of it has a mass of 20g.

Solution:

$$n = \frac{m}{M} \Rightarrow M = \frac{m}{n} = \frac{20g}{0.25mol} = 80g/mol$$

107. Calculate the percentage composition by mass of element in $Na_2Cr_2O_7$.

Solution:

$Na_2Cr_2O_7$

$Na: 2\times23g = 46g$

$Cr: 2\times52g = 104g$

$O: 7\times16g = 112g$

mass of $Na_2Cr_2O_7 = 262g$

$$\%Na = \frac{\text{mass Na}}{\text{mass } Na_2Cr_2O_7}\times100$$

$$= \frac{46g}{262g}\times100 = 17.6\%$$

$$\%Cr = \frac{\text{mass Cr}}{\text{mass } Na_2Cr_2O_7}\times100$$

$$= \frac{104g}{262g}\times100 = 39.7\%$$

$$\%O = 100 - (\%Na + \%Cr)$$

$$= 100 - (17.6 + 39.7) = 42.7\%$$

108. Describe the reaction below stoichiometrically.

$$2NH_3(g) + 3O_2(g) \rightarrow 2N_2(g) + 6H_2O(g)$$

Solution:

Based on mole - ratio

- 1mole of N_2 is produced per mole of NH_3
- 2mole of H_2O is produced per mole of O_2

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- 3mole of H_2O is produced per mole of NH_3.
- 2mole of N_2 is produced for every 3mole of O_2.

109. 2 grams of a mixture of Cu_2O and CuO was reduced to 0.8g of Cu. What was the weight in grams of CuO in the original sample?

Solution:

Mass of $(Cu_2O + CuO) = 224g$

Formula mass of Cu_2O and $CuO = 224g$

$$\text{mass of CuO in the original sample} = \frac{\text{formal mass of CuO}}{\text{formal mass of } Cu_2O \text{ and CuO}} \times \text{mass of mixture}$$

$$= \frac{80g}{224g} \times 2g = 0.714g$$

110. How many litters will 3.5gof hydrogen, N_2 occupy at STP?

Solution:

$$1\text{mol } N_2 = 28g\ N_2 = 22.4L\ N_2 \text{ at STP}$$

$$3.5g\ N_2 = ?\ x$$

$$x = \frac{3.5g\ N_2 \times 22.4L}{28g} = 2.8L$$

$\therefore$ 3.5g of N_2 will occupy 2.8L at STP.

111. How many liters of CO_2 at STP will be produced according to the reaction: $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$ from 11.2L of O_2 and excess carbon monoxide, CO?

Solve by your own! Answer: 22.4L of CO_2

112. What is the molecular mass of a gas if 150mL of the gas measured at STP, weighs 0.228g?

Solution:

1 mole of gas = molar mass of gas = 22400mL

$$M = 22400mL$$

$$0.228g = 150mL$$

$$\Rightarrow M = \frac{0.228g \times 22400mL}{150mL} = 34g$$

113. How many liters of ammonia (NH_3) is formed when 36L hydrogen reacts with 13L nitrogen gas at STP?

Solution:

$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

Identify the limiting reactant by dividing to its coefficient.

$$V_{N_2} = \frac{13L}{1} = 13L$$

$$V_{H_2} = \frac{36L}{3} = 12L$$

$\therefore$ H_2 is limiting reactant.

$$N_2(g) + \underset{3}{\overset{36L}{3H_2(g)}} \longrightarrow \underset{2}{\overset{x}{2NH_3(g)}}$$

$$\frac{36L}{3} = \frac{x}{2}, \quad x = V_{NH_3} = 24L$$

114. Calculate the mass percent of $MgSO_4$ in $MgSO_4 . 7H_2O$

Solution:

Mass of $MgSO_4 = 1\times24 + 1\times32 + 4\times16 = 120g$

Mass of $MgSO_4.7H_2O = 120 + 7\times18 = 246g$

$$\%MgSO_4 = \frac{120g}{246g} \times 100 = 48.78\%$$

115. Find the formula of a hydrate of with compositions: Ca = 18.2%, Cl = 32.42% and H_2O = 49.31%.

Answer: $CaCl_2.6H_2O$

116. Find the molar mass of a gas when the density is given as 1.25g/L.

Solution:

Molar mass = density × molar volume

$$= 1.25\frac{g}{\cancel{L}} \times 22.4\frac{\cancel{L}}{mol} = 28.0 g/mol$$

117. One mole of DAP, $(NH_4)_2HPO_4$, has mass of 131g. What is the percentage by mass of nitrogen in DAP?

Answer:21.21%

118. One atom of an element has a mass of 4.48×10^{-23}g. What is the approximate molar mass of the element?

Solution:

$M = 6.02\times10^{23}$ atoms/mol

$4.48\times10^{-23} g = 1 atom$

$\Rightarrow M = 4.48\times10^{-23} g/atom \times 6.02\times10^{23}$ atoms/mol $= 27 g/mol$

119. How many moles of oxygen atoms are there in 200g of $CaCO_3$?

Solution:

100g of $CaCO_3$ = 3mol O

200g of $CaCO_3$ = ?

$$\Rightarrow n(O) = \frac{200g \times 3mol}{100g} = 6mol$$

120. 3.6g of magnesium metal reacted completely with pure nitrogen in a closed container to give 5g of a compound. What is the formula of this compound?

Solution:

mass of nitrogen = mass of compound - mass of Mg = 5g - 3.6g = 1.4g

S	O
$\frac{50}{50}=1$	$\frac{50}{50}=1$
$\frac{50}{50}=1$	$\frac{75}{50}=1.5$

The ratio of oxygen in thiese compounds is 2:3

121. Given two hypothetical compounds A and B that contain sulphur and oxygen according to the information given in the table below

Compound	Mass of S in grams	Mass of O in gram
A	50	50
B	50	75

What is the ratio of the masses of oxygen in these compounds?

Solution:

S	O
$\frac{50}{50}=1$	$\frac{50}{50}=1$
$\frac{50}{50}=1$	$\frac{75}{50}=1.5$

The ratio of oxygen in thiese compounds is 2:3

122. Given the unbalanced equation representing the reaction: $P + O_2 \rightarrow P_2O_5$ How many grams of phosphorus should be burnt to produce 14 .2g of phosphorus pentoxide.

Answer: 6.2g of phosphorus

123. Given the following balanced equation representing the combustion of methane: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$. How many litter of oxygen are required to produce 2 litters of carbon dioxide?

Answer: 4L of O_2

Supplementary Exercise

I. True/False items

1. All double displacement reactions are non - redox reactions.
2. All disproportionation reactions are redox reactions.
3. Oxidation is decrease in oxidation state due to lose of electrons.
4. Formula mass and molar mass have the same numerical value but different units.
5. Chemical change are easily reversed and occur due to chemical reaction
6. Reactants are real materials usually written to the left side of the arrow.
7. Unbalanced chemical equation can be used to describe the relationships

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between the amount of reactants and products.

8. Limiting reactant determines the amount of products formed.
9. In single displacement reaction, more reactive element displaces less reactive element.
10. Oxidizing agent causes oxidation by losing electrons.

II. Matching Items

A	B
11. Combination reaction	A. 6.02×10^{23}
12. Decomposition reaction	B. 100
13. Single displacement reaction	C. 22400mL
14. Double displacement reaction	D. Analysis
15. Avogadro's number	E. Metathetical reaction
16. Molar Volume	F. Synthesis
17. Sum of percentage composition	G. Formula mass in gram
18. Molar mass	H. Actual formula
19. Empirical formula	I. Simplest formula
20. Molecular formula	J. Substitution reaction
	K. 11.2L
	L. 3.01×10^{23}

III. Multiple Choice Items

21. All of the following are chemical changes. **EXCEPT**
   A. Souring of milk
   B. Fermentation of alcohol
   C. Heating of sugar
   D. Sublimation of iodine
22. Which of the following represents the chemical equation for burning of magnesium ribbon?
   A. $Mg + O \rightarrow MgO$
   B. $2Mg + O_2 \rightarrow 2MgO$
   C. $Mn + O \rightarrow MnO$
   D. $4Mn + 3O_2 \rightarrow 2Mn_2O_3$
23. Which of the following is a single displacement reaction?
   A. The reaction between acid and base
   B. Heating of metal hydrogen carbonate
   C. The reaction between active metal with water
   D. The reaction between metal and non - metal

24. The oxidation state of Mn in $KMnO_4$ is
   A. +7 B. +6 C. +5 D. +4
25. The formula mass of $Al_2(SO_4)_3$ is
   A. 342 B. 171 C. 180 D. 310
26. The coefficient of hydrochloric acid, HCl in the reaction below after balancing it is
   $KMnO_4 + HCl \rightarrow KCl + MnCl_2 + Cl_2 + H_2O$
   A. 5 B. 16 C. 8 D. 2
27. What is the reducing agent in the reaction below?
   $2HNO_3 + 3SO_2 + 2H_2O \rightarrow 3H_2SO_4 + 2NO$
   A. HNO_3 B. SO_2 C. H_2O D. H_2SO_4
28. How many grams of phosphorus are there in 20g of $Ca_3(PO_4)_2$?
   A. 62g B. 4g C. 31g D. 15.5g
29. In which of the following species does nitrogen have the lowest oxidation number?
   A. N_2H_4 B. N_2O_4 C. NH_3 D. N_2O_5
30. Which of the following is wrong about 2 moles of $(NH_4)_2Cr_2O_7$? It has
   A. 4moles of nitrogen atoms
   B. 16moles of hydrogen atoms
   C. 14moles of oxygen atoms
   D. 2moles of chromium atoms
31. Given the equation: $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$. At STP, how many liters of hydrogen will react with 20 liters of oxygen?
   A. 40L B. 20L C. 10L D. 22.4L
32. Which of the following statements is not true for the following reaction?
   $Fe^{2+} + MnO_4^- + H^+ \longrightarrow Fe^{3+} + Mn^{2+} + H_2O$
   A. Fe^{2+} loses electron
   B. Fe^{2+} is reducing agent
   C. MnO_4^- is oxidizing agent
   D. H^+ gains electrons
33. How many moles of aluminium oxide, Al_2O_3 can be formed when a mixture of 0.36mol Al and 0.36mol of O_2 gas is ignited?
   A. 0.72mol B. 0.12mol C. 0.46mol D. 0.18mol
34. If 100mL of SO_2 and 100mL of O_2 are reacted to fromSO_3, how many mL of excess reacted will left at the end of the reaction?
   A. 50ml of O_2
   B. 50ml of SO_2

C. 25ml of O_2

35. The percentage of water in $Na_2CO_3 . xH_2O$ is 62.9%. What is the value of x?

A. 7 B. 10 C. 5 D. 6

36. How many liters of hydrogen gas are necessary to react with sufficient chlorine to produce 16 liters of hydrogen chloride gas?

A. 16L B. 8L C. 12L D. 6L

37. What mass of oxygen can be obtained from heating 100g of potassium chloride? $2KClO_3 \rightarrow 2KCl + 3O_2$

A. 245g B. 96g C. 48g D. 39g

38. 2grams of hydrogen at STP is not equal to

A. 1 mole of H_2
B. 22.4L of H_2
C. 6.02×10^{23} molecules of H_2
D. 6.02×10^{23} atoms of H

39. Quicklime, CaO is to be prepared by roasting 2kg of limestone, $CaCO_3$. The actual yield of CaO is 1.05kg. What is the percentage yield?

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

A. 93.7% B. 67.8% C. 88.6% D. 79.3%

40. When a sample of a gas was analysed it was found to contain 85.72% carbon and 14.28% hydrogen if the relative mass of the gas is 28. What is the molecular formula of the gas?

A. CH B. C_2H_2 C. C_2H_4 D. C_2H_6

IV. Short Answer Items

41. Classify the following as physical or chemical properties of substances.

a. Odour
b. Fermentation
c. Density
d. Rusting
e. Respiration
f. Malleability
g. Contraction
h. Freezing point
i. Reactivity
j. Hardness
k. Flammability

42. Convert the following Chemical equation into word equation?

a. $Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$ b. $P_4 + 5O_2 \rightarrow 2P_2O_5$

43. Change the following into chemical equation

a. Calcium hydrogen + phosphoric acid → Calcium phosphate + water
b. Magnesium + nitrogen → Magnesium nitride.

44. _____ deal with quantitative relative relationship between reactants and products in a balanced chemical equation.

45. If the same substance is oxidized and reduced at the same time, the reaction is _____

V. Workout Questions

46. Balance the following the following chemical equations.

a. $MnO_2 + KOH + O_2 \rightarrow KMNO_4 + H_2O$
b. $Fe_2O_3 + C \rightarrow Fe + CO$
c. $(NH_4)_2CO_3 + H_2O \rightarrow NH_4OH + H_2CO_3$

47. How many moles of water is formed from the complete combustion of 5 moles of hexane (C_6H_{14})?

48. Make the following conversions.

a. 5.8×10^{24} formula units of $CaCO_3$ to moles and grams
b. 20g of HCl to moles and molecules
c. 0.25mol caffeine, $C_8H_{10}O_2N_4$, to total number of atoms and mass in gram.

49. When 56g of calcium and excess amount nitrogen gas undergo a reaction that has a 90%-yield, what is the actual yield?

50. In an experiment to prepare bromo benzene according to the equation,

$$C_6H_6 + Br_2 \rightarrow C_6H_5Br + HBr.$$

A student reacted 20g of C_6H_6 with 0.310mol of Br_2

a. What is the limiting reactant
b. How many grams of the excess reactant last.

Answers

Part I

1. True
2. True
3. False
4. True
5. False
6. True
7. False
8. True
9. True
10. False

Part II

11. F
12. D
13. J
14. E
15. A
16. C
17. B
18. G
19. I
20. H

Part III

21.D	24.A	27.B	30.D	33.D	36.B	39.A
22.B	25.A	28.B	31.A	34.A	37.D	40.C
23.C	26.B	29.C	32.D	35.B	38.D	

Part IV

41. a, c, f, g, h and j are physical properties.
    b, d, e, i, and k, are chemical properties.

42. a. Iron + sulphuric acid → Iron(11)sulphate + hydrogen
    b. Phosphorus + oxygen → phosphors pentoxide.

43. a. $3Ca(OH)_2 + 2H_3PO_4 \rightarrow Ca_3(PO_4)_2 + 6H_2O$
    b. $3Mg + N_2 \rightarrow Mg_3N_2$

44. Stoichiometry

45. Disproportionation reaction

Part V.

46.
    a. $4MnO_2 + 4KOH + 3O_2 \rightarrow 4KMnO_4 + 2H_2O$
    b. $Fe_2O_3 + 3C \rightarrow 2Fe + 3CO$
    c. $(NH_4)_2CO_3 + 2H_2O \rightarrow 2NH_4OH + H_2CO_3$

47. $\overset{5mol}{2C_6H_{14}} + 19O_2 \longrightarrow 12CO_2 + 14\overset{x}{H_2O}$

$\frac{5mol}{2} = \frac{x}{14}$, $x = 35$ moles of H_2O are formed.

48. a) 1mole $CaCO_3 = 6.02\times10^{23}$ formula units
? $= 5.8\times10^{24}$ formula units
$\Rightarrow 5.8\times10^{24}$ formula units = 9.6 moles of $CaCO_3$
100g $CaCO_3 = 6.02\times10^{23}$ formula units
? $= 5.8\times10^{24}$ formula units
5.8×10^{24} formula units = 960g of $CaCO_3$
b) 20g HCl = 0.55mol of HCl
$= 3.3\times10^{23}$ molecule of HCl
c) 0.25mol $C_8H_{10}O_2N_4 = 3.612\times10^{24}$ atoms

$= 48.5g$ of $C_8H_{10}O_2N_4$

49. $\overset{56g}{3Ca} + N_2 \longrightarrow \overset{Ty}{Ca_3N_2}$
$\quad\quad 3\times40g \quad\quad\quad 1\times148g$

$\frac{56g}{3\times40g} = \frac{Ty}{1\times148g} \Rightarrow Ty = 69g.$

$Ay = \frac{\%yield \times Ty}{100} = \frac{90\times69g}{100} = 60.1g$

50. a) C_6H_6
b) 8.57g of Br_2

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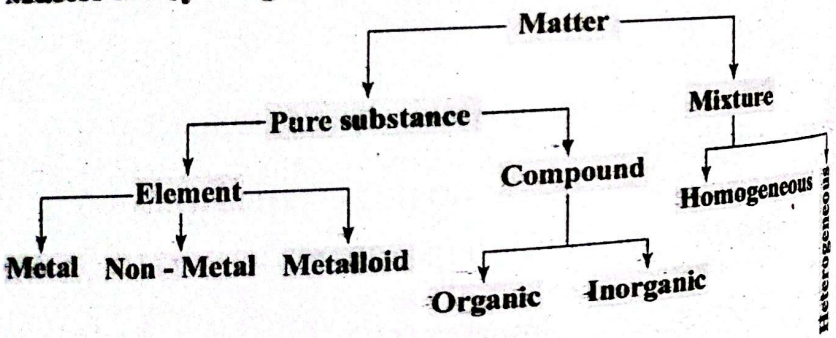
UNIT 7

Solution

Matter is anything that has mass and occupies space.



Mixtures

- ☑ They result from combination of two or more pure substances in **variable proportions.**
- ☑ **Substances** in a mixture have not undergone a chemical reaction.
- ☑ It is possible to separate the components of a mixture provided there is a suitable difference in their physical properties.

7.1 Types of Mixtures

Based on the number of phases they possess, mixtures are classified into two. These are:

1. Heterogeneous mixture
2. Homogeneous mixture

I. Heterogeneous Mixture

It is a type of mixture that has different phases.

Phase: Uniform part of a system that is in contact with other parts of the system but separated from them by a well defined boundary.

Example of heterogeneous mixture:

- ➥ Blood
- ➥ Milk
- ➥ Coffee
- ➥ Soil
- ➥ Sand and water
- ➥ Oil and water
- ➥ Dough etc.

Suspensions are heterogeneous mixtures in which particles settle down after a given period of time.

Example

- ➥ Milk of magnesia ($Mg(OH)_2$)
- ➥ A mixture of sand and water
- ➥ A mixture of mud and water
- ➥ A mixture of chalk and water

All suspensions are heterogeneous mixtures but all heterogeneous mixture are not necessariry suspensions.

Colloids are mixtures containing small particles that are small enough not to settle downwards.

- ➥ They are heterogeneous mixture.
- ➥ They scatter abeam of light passing through them.

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Tyndall Effect is the phenomenon in which the particle in a colliod scatter the beam of light.

Ostwald regarded colloidal systems as heterogeneous in character, consisting of two phases.

i. Disperse phase
ii. Dispersion medium

Disperse phase refers to the phase forming the particles.

Dispersion medium is the medium in which dispersion of the particles takes place.

Table 7.1 Types of Colloidal Dispersions

Dispersion medium	Dispersion phase	Name	Examples
Gas	Gas	-	Unknown
	Liquid	Aerosol	Fogs, mists, clouds
	Solid	Aerosol	Smokes
Liquid	Gas	Foam	Whipped cream
	Liquid	Emulsion	Dispersion of oil in water or water in oil
	Solid	Sols	AgCl, Au, AS$_2$S$_3$ or S in water
Solid	Gas	-	Solid foam
	Liquid	Gels	Jellies minerals with liquid butter,cheese
	Solid	-	Glasses coloured with dispersed metals. ruby glass

Dispersion of solids in liquids can be roughly divided into two categories.

i. **Lyophobic (solvent hating) sols** have attraction between the disperse phase and dispersion medium.
ii. **Lyophilic (solvent loving) sols** are systems in which the disperse phase shows some definite affinity for the medium.

Table 7.2 Characteristics of Lyophobic and Lyophilic Colloids

Lyophobic Colloids	Lyophilic Colloids
➥ Generally unstable,	➥ Very stable
➥ Easly coagulated on adding electrolytes	➥ Not easity coagulated by electrolytes
➥ Irreveresible in character	➥ Reversible
➥ Prepared by indirect method	➥ Usually prepared by simple solution method
➥ Viscosity of colloidal dispersion is same as that of the solvent	➥ They are highly viscous systems
➥ They are generally inorganic materials i.e metals, sulphides and oxide sols	➥ Mostly organic materials e.g starch, proteins, gums etc
➥ Carry electric charge	➥ Carry little or no charge

2. Homogeneous Mixture

It is a type of mixture that has one phase or uniform composition

Example

- Pure air
- Sea water
- Alloys (steel, brass, bronze, solder etc)
- Soft drinks (coca - cola, Pepsi, spirit etc)
- Distilled alcoholic beverages (Whisky, gin katicala or Areki etc)
- Vinegar or Acheto
- Salt solution
- Sugar solution etc

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Illustrative Example

1. Which one is not a homogeneous mixture?

A) Milk C) Whisky

B) Brass D) Coca-cola drink

Solution

Milk is heterogeneous mixture.

Answer: A

2. Which one of the following is pure suspension?

A) Smoke C) Fog

B) Cloud D) Milk of magnesia

Answer: D

3. Which one of the following is pure substance?

A) Air B) Food C) Shampoo D) Salt

Solution

Pure substances are either elements or compounds. Salt is a compound but the rest are mixtures

Answer: D

4. The type of mixture with the same composition throughout is:

A) Suspension C) Homogeneous

B) Colloid D) Heterogeneous

Solution

Suspension and colloid are heterogeneous mixtures.

Answer: C

5. Classify the following as pure substance, homogeneous or heterogeneous mixture

a) Air c) Gold e) Blood g) Natural gas

b) Water d) Mineral f) Fruit punch h) Sea water

Solution

- a, g and h are homogeneous mixture
- b and c are pure substances
- d,e and f are heterogeneous mixture

6. List some properties of colloids

Solution

- They are heterogeneous mixture
- The particles in a colloid is smaller in size than suspension but bigger than solution
- They are stable
- They scatter a beam of light passing through them.

7. The phenomenon in which the particle in colloid scatter the beams of light that are directed at them and make the path of the light beam visible is known as _____

Answer: Tyndall effect

8. Define a) micelles b) solution

Solution

a) **Micelles** are colloidal - sited particles resulted when molecules or ions of substances that have hydrophobic and hydrophilic ends are dispersed in water. Soaps and detergents are such substances with hydrophobic (water hating) and hydrophilic (water loving) tail and head respectively.

b) It is homogeneous mixture

It has two components

i. **Solute** - usually exists in smaller amount
   - undergoes phase change
   - is dissolved component

ii. **Solvent** - usually occurs in larger quantity
   - does not undergo phase change
   - affects the state of a solution

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Example

In sugar solution

Solute = sugar

Solvent = water

Types of Solution

Based on the state of solute and solvent, solutions can be classified into six basic types.

Table 7.3 Types of Solutions and Examples

No	State of solute	State of solvent	State of solution	Type of solution	Example
1	Gas	Gas	Gas	Gas - Gas	Air, natural gas
2	Gas	Liquid	Liquid	Gas - liquid	Coca - cola Ambo water
3	Liquid	Liquid	Liquid	Liquid - liquid	Whisky Areki
4	Liquid	Solid	Solid	Liquid - solid	Amalgam (mercury silver)
5	Solid	Liquid	Liquid	Solid - liquid	Salt solution, Sugar solution
6	Solid	Solid	Solid	Solid - solid	Brass(Cu+Zn) Bronze (Cu + Sn)

Based on the state of solutions, we can classify solutions into three. These are:

A) Gaseous solution is solution with gaseous state

Example

- Air ($N_2 + O_2 + Ar$ + other gases)
- Natural gas ($CH_4 + C_2H_6 + C_3H_8 + C_4H_{10}$)

- Water gas ($CO + H_2$)
- Producer gas ($CO + N_2$)

B) Liquid solution is a solution with liquid state

Example

- Ethanol in water
- Oil in benzene
- Distilled alcoholic beverages
- Sugar solution

C) Solid solution is a solution that has solid state

Example

- Alloys (bronze, amalgam etc)

Illustrative Examples

9. What is the solvent in 70% alcohol solution?

A) Water B) Alcohol C) Sugar D) Kerosene

Solution

Even though the amount of alcohol is higher than water, water is the solvent since it is universal polar solvent.

Answer: A

10. What kind of solution forms when gasoline evaporates in air?

A) Gas in liquid solution C) Liquid in liquid solution

B) Gas in gas solution D) Liquid in gas solution

Solution

Gasoline is converted into gaseous state since it is evaporated. Thus, it is gas - gas solution.

Answer: B

11. Which of the following types of solutions are possible?

I. Solid dissolved in a liquid

II. Gas dissolved in a liquid

III. Gas dissolved in a gas

IV. Solid dissolved in a solid

A) I and II
C) I, II, III and IV
B) I, II and IV
D) I

Answer: C

12. What are the possible physical states of solute and solvent in

a) Gaseous solution?
b) Liquid solution?
c) Solid solution?

Solution

State of solute	State of solvent
a) Gas	Gas
b) Gas, liquid, solid	liquid
c) Liquid, solid	Solid

13. Identify the solute and solvent in the solutions given below

a) A syrup made of 60% sugar and 40% water
b) Red brass, 90% copper and 10% zinc
c) A 1: 1 mixture of benzene and toluene
d) 40g salt and 20g of water

Solution

Solute	Solvent
a) Sugar	Water
b) Zinc	Copper
c) Benzene /toluence	toluene /benzene
d) Salt	Water

14. How could you differentiate solution from colloid?

Solution

Solution does not scatter a beam of light passing throught but s colloid does.

7.2 The Solution Process

It is the process of forming a solution by dissolving the solute completly in a solvent

It involves three steps.

Step 1: Separating the solute particle

- It is breaking of inter - particle force that holds the solute particles together.

  For ionic solute, $\Delta H_1 = \Delta H_{solute} = \Delta H_{LE}$

- The energy required to break an ionic solid into separated ions is know as lattice energy (ΔH_{LE})

$\Delta H_{sol} = \Delta H_{LE} + \Delta H_{solvent} + \Delta H_{mix} = \Delta H_{LE} + H_{hyd}$

It is an endothermic step (absorbs energy) ($\Delta H_{solute} > 0$)

Step 2: separating solvent particles.

- It is breaking of solvent - solvent inter - particle force to give space for the incoming solute particles
- It is also an endothermic step ($\Delta H_{solvent} > 0$)

Step 3: The solvent molecules surround the solute particles. This process is called **solvation** and if the solvent is water, it is termed as **hydration**.

- It involves the formation of soluter - solvent interactions.
- Most of the time it is an exothermic step (releases energy) ($\Delta H_{mix} < 0$)

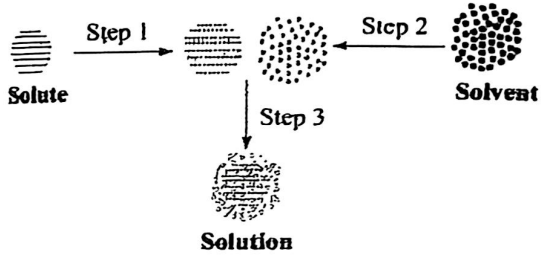

Fig 7.1 The Solution Process Occurs in Three Distinct Steps

IV. Solid dissolved in a solid

A) I and II
B) I, II and IV
C) I, II, III and IV
D) I

Answer: C

12. What are the possible physical states of solute and solvent in
a) Gaseous solution?
b) Liquid solution?
c) Solid solution?

Solution

State of solute	State of solvent
a) Gas	Gas
b) Gas, liquid, solid	liquid
c) Liquid, solid	Solid

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d) 40g salt and 20g of water

Solution

Solute	Solvent
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c) Benzene /toluence	toluene /benzene
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$\Delta H_{sol} = \Delta H_{LE} + \Delta H_{solvent} + \Delta H_{mix} = \Delta H_{LE} + H_{hyd}$

It is an endothermic step (absorbs energy) ($\Delta H_{solute} > 0$)

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- It is also an endothermic step ($\Delta H_{solvent} > 0$)

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- It involves the formation of soluter - solvent interactions.
- Most of the time it is an exothermic step (releases energy) ($\Delta H_{mix} < 0$)

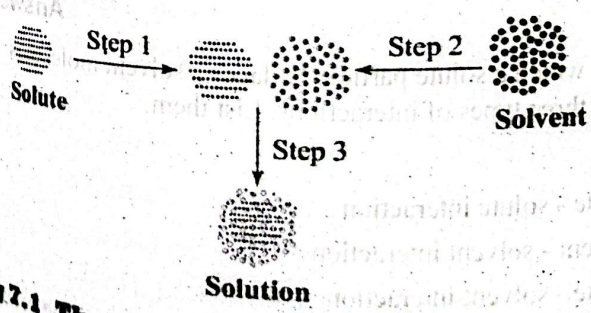

Fig 7.1 The Solution Process Occurs in Three Distinct Steps

Note: For gas - gas solution step1 and step 2 are not required because gaseous molecules are already separated.

☑ Formation of solution becomes easier if solute - solvent interaction dominates both solute - solute and solvent - solvent interactions.

☑ The solute - solvent interaction to give a solution is a physical process.

☑ The reaction between strong acid and strong base to give a salt solution is a chemical process.

Illustrative Example

15. The process of solute particles being surrounded by solvent particles is known as

A) Saturation C) Solvation

B) Agglomeration D) Dehydration

Answer: C

16. What is the exothermic step in a solution of $CaCl_2$ in water?

A) Separation of the salt in to ions in gaseous state

B) Separation of water molecules.

C) Hydration (solvation) of the ions from the salt.

D) Separation of the salt in to ions in gaseous state and separation of the water molecules.

Solution

Hydration or solvtion is an exothermic step.

Answer: C

17. The ease in which a solute particle replaces a solvent molecules depends on three types of interactions. List them

Solution

➥ Solute - solute interaction

➥ Solvent - solvent interaction

➥ Solute - solvent interaction

18. When the solute and solvent interact to form a solution, the process is

Answer: Physical process.

Energy Changes in the Solution Process

Heat of solution (ΔH_{sol}) is the amount of energy absorbed or released during the solution process.

☑ It is the sum of the heat changes associated with the three steps in the solution process.

$$\Delta H_{sol} = \Delta H_{solute} + \Delta H_{solvent} + \Delta H_{mix}$$

ΔH_{solute} = heat of solute

$\Delta H_{solvent}$ = heat of solvent

ΔH_{mix} = heat of mixture of solute and solvent.

☑ **Enthalpy change (ΔH)** is heat transferred at constant pressure (q_p)

$$\Delta H = q_p$$

Thus, $\boxed{\Delta H_{sol} = q = mC_g\Delta T}$

C_g - specific heat of solution

m - mass of solution

$\Delta T = T_i - T_f$

☑ The unit of heat of solution is joule (J) or kilo joule (kJ)

$\Delta H_{sol} > 0$, endothermic mixing

$\Delta H_{sol} < 0$, exothermic mixing

$\Delta H_{sol} = 0$, ideal solution

☑ **Molar heat of solution (ΔH_{mol}) of solute** is the heat absorbed or released when one mole of a solute is dissolved in water.

$$\boxed{\Delta H_{mol} = \frac{q}{n_{solute}}}$$

➥ n_{solute} - number of mole of solute

➥ The unit of ΔH_{mol} is kilo joule per mol (kJ/ mol)

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Example

When 10g of a salt MX is dissolved in 90g of water, the temperature of water rises by 4°C. Calculate

a) ΔHsol or q

b) ΔHmol

(Cg = 4J/g °C, molar mass of MX = 40 g/mol)

Solution

a) Mass of solution = mass of + mass of
solute solvent
= 10g + 90g = 100g

A rise in temperature means $T_f > T_i$.

$\Rightarrow \Delta T = T_i - T_f = -4^oC$

$q = m\ Cg\ \Delta T$

$$q = 100g \times 4\frac{J}{C^o.g} \times (-4^oC) = -1600J$$

b) $$\Delta Hmol = \frac{q}{n_{solute}}$$

$$nsolute = \frac{m}{M} = \frac{10g}{40g/mol} = 0.25mol$$

$$\Rightarrow \Delta H_{smol} = \frac{-1600J}{0.25mol} = -6400J/mol$$

Note

i. For gaseous solute and solvent
$\Delta H_{sol} = \Delta H_3$ because $\Delta H_1 = 0,\ \Delta H_2 = 0$

ii. Heat of solvation (ΔH_{solv}) is the amount of energy released when solvent molecules surround solute particles.

iii. Heat of hydration (ΔH_{hyd}) is the amount of energy released when water molecules surround solute particles.
$\Delta H_{hyd} = \Delta H_{solvent} + \Delta H_{mix}$

c) $\Delta H_{sol} = 0$, Ideal solution

d) $\Delta H_{sol} = \Delta H_{mix}$ for any solute and solvent

Illustrative Example

19. When pure methanol is mixed with water, the solution becomes warm. Is this an ideal solution?

A) Yes, for ideal solution, $\Delta H_{sol} = 0$

B) No, for ideal solution, $\Delta H_{sol} = 0$

C) Yes, for ideal solution, $\Delta H_{sol} > 0$

D) No, for ideal solution , $\Delta H_{sol} < 0$

Solution

The solution is warm, meaning it is formed by exothermic mixing. For ideal solution, $\Delta H_{sol} = 0$

Answer: B

20. When you touch a device that contains a solution, you feel cold. The given solution is prepared:

A) By endothermic mixing C) Ideally

B) By exothermic mixing D) None

Solution

We can easily identify by touching a given solution whether it is formed by endothermic or exothermic mixing. If we feel hot, it is prepared by exothermic mixing but if we feel cold, it is formed by endothermic mixing.

Answer: A

21. Which one is not correct about heat of solution?

A) ΔHsol > 0, endothermic mixing

B) ΔHsol < 0, exothermic mixing

C) ΔHsol = 0 , ideal solution

D) None

Answer D

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22. Gases mix exothermically. Why?

Solution

Reason: $\Delta H_{solute} = H_{solvent} = 0$ because the gaseous molecules are already separated and enthalpy change of the mixture (ΔH_{mix}) is less than zero.

$\Rightarrow \Delta H_{sol} = \Delta H_{mix} < 0$. That is why gases mix exothermically.

The Rate of Dissolution

☑ The speed at which the solute enters into a solution is called rate of dissolution.

☑ The rate of dissolution is mainly affected by inter - particle forces (inter-molecular forces)

☑ The main inter - particle forces responsible for the solubility of solutes are

1. Ion - dipole Force

➥ It is involved when ionic solutes dissolve in water.
Example: sodium chloride in water.

2. Hydrogen Bonding

➥ It is the principal force in solutions of polar H - F, O - H and N - H bonding containing inorganic, organic and biological compounds in water

Example: Ethanol in water.

3. Dipole - Dipole Force

➥ It allows the solubility of polar molecules in the absence of hydrogen bond in polar solvent.

Example: Hydrogen chloride in water

4. Dispersion Force (London force)

➥ It contributes the solubility of non - polar solute in non - polar solvent.

Example: Benzene in carbon tetra chloride

5. Ion - induced Dipole Forces

➥ It is one type of charge - induced dipole force, rely on polarizability.

➥ It arises when an ion's charge distorts the electron cloud of a nearby non - polar molecule.

Example: The binding of Fe^{2+} ion in hemoglobin to an O_2 molecule entering a red - blood cell.

➥ It arises when a polar molecule distorts the electron cloud of non - polar molecule.

Rate of dissolution is also affected by

➥ Temperature

➥ Surface area of solid solute

➥ Pressure of gaseous solute

Solutions of Liquids in Liquids

The solubility of solute in a solvent follows the "**like dissolves like principle**". That means polar dissolves polar and non - polar dissolves non - polar.

Example

➥ Water and ethanol are miscible due to hydrogen bonding and both are polar.

➥ Benzene and toluene are miscible because of London force and both are non - polar.

➥ Oil and water are immiscible because oil is non - polar but water is polar.

Solution of Solids in Liquids

Polar solids dissolve in polar solvent and non - polar solids dissolve in non - polar solvent.

Example

➥ Iodine dissolve in benzene due to London dispersion force.

➥ Sugar is soluble in water because of hydrogen bond.

➥ Sodium chloride is soluble in water due to ion - dipole force.

Note: The solubility of ionic compounds in water is affected by hydration energy and lattice energy.

☑ The greater the hydration energy, the more likely that ionic solid will dissolve in water.

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☑ The lower the lattice energy, the more soluble is the ionic compound in water

☑ Lattice energy is affected by the size of the ions. Their relationship is inverse.

Example

Arrange the following ionic compounds in decreasing order of solubility in water

NaCl, NaF, NaI , NaBr

Solution

NaI > NaBr > NaCl > NaF

Reason: size of I^- > Br^- > Cl^- > F^-

The larger the size, the smaller is the lattice energy. As a result, the higher is the solubility.

Illustrative Examples

23. The rate of dissolution is mainly affected by

A) Temperature C) Surface area

B) Pressure D) Interparticle force

Answer: D

24. What is the most important type of soluter - solvent interaction in a solution of ammonia in water?

A) Ion - dipole C) Dipole - dipole force

B) Hydrogen bonding D) London dispersion forces

Solution

Both ammonia (NH_3) and water (H_2O) have hydrogen bonding

Answer: B

25. Which of the following is most likely to be miscible in carbon tetrachloride, CCl_4?

A) Na_2CO_3 B) CH_3F C) $(CH_2)_2O_2$ D) I_2

Solution

CCl_4 is non-polar solvent that dissolves non - polar solute.

Answer: D

26. Identify the solute - solvent interaction responsible for the solubility of

a) HBr in H_2O c) HF on H_2O

b) $AgNO_3$ in H_2O d) Br_2 in C_6H_6

Solution

a) Dipole - dipole force c) Hydrogen bonding

b) Ion - dipole force d) London dispersion force

27. Sodium carbonate (Na_2CO_3) is soluble but calcium carbonate ($CaCO_3$) is insoluble in water why?

Solution

Reason: Ca CO_3 has higher lattice energy than Na_2CO_3.

28. Arrange the following ionic compounds in increasing order of solubility.

LiCl, KCl, RbCl, Na Cl

Solution

Li Cl < NaCl < KCl < RbCl

29. What are Van der waals forces?

Solution

They dipole - dipole, induced dipole and dispersion forces.

30. Explain the solubility rules of ionic solids.

Solution

- All common salts of alkali metals (Li, Na,K), ammonium ion (NH^+_4), acetate ion (CH_3COO^-), nitrate ion (NO^-_3) and most perchlorate (ClO^-_4) are soluble in water.
- All common halides (Cl^-, Br^-, I^-) except that of Ag^+, Pb^{2+} and Hg^{2+}_2 all fluorides except that of Pb^{2+} and group IIA are soluble.
- All compounds containing sulphate ion $\left(SO^{2-}_4\right)$ except that of Ba, Sr, Ca, Ag and Hg are soluble

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- All carbonates (CO_3^{2-}) phosphates (PO_4^{3-}), sulphides (S^{2-}), are insoluble except of hose of alkalimetals and ammonium ion.

31. Arrange the following in increasing order of induced dipole interactions.
He, Ne, Ar, Kr, Xe

Solution

The larger the size, the higher is the induced dipole interactions since it is easily polarized

$\Rightarrow$ He < Ne < Ar < Kr < Xe

7.3 Solution Formation and Equilibrium

Based on equilibrium, solution can be classified into three. These are

1. Saturated solution
2. Unsaturated solution
3. Supersaturated solution

Table 7.4 Difference Between Saturated, Unsaturated and Supersaturated Solution

Saturated solution	Unsaturated solution	Supersaturated solution
➥ Contains maximum amount of dissolved solute at a given temperature ➥ Always occurs at equilibrium ➥ Size of an added crystal remains same	➥ Has less solute than saturated solution ➥ Dissolves additional solute at a given condition. ➥ Size of an added crystal decreases	➥ Has more solute than saturated solution ➥ Develops a crystal easily and is unstable. ➥ Size of crystal grows bigger when added to it

Effect of Temperature on Solubility

Solubility is the maximum amount of solute that will dissolve in a given amount of solvent at a given temperature to form stable solution.

☑ For endothermic mixing ($\Delta H_{sol} > 0$), increasing temperature increases solubility. and vice - versal

☑ For exothermic mixing ($\Delta H_{sol} < 0$), increasing temperature decreases solubility.

Effect of Pressure on Solubility

☑ Pressure has little effect on solubility of solid and liquid solutes.

☑ Pressure has effect on solubility of gaseous solute.

☑ As pressure of a gas increases, its solubility also increases. This was explained by Henry's law.

☑ **Henry's law** states that the solubility of a gas (C_g) is directly proportional to the partial pressure of the gas (P_g) above the solution.

$$C_g \ \alpha \ P_g$$
$$C_g = K_H P_g$$

K_H - Henry's constant

Example

The partial pressure of carbon dioxide gas inside a bottle of coca - cola is 4atm at 25°C. What is the solubility of CO_2 in the soft drink?

(K_H for $CO_2 = 3.3 \times 10^{-2}$ mol/l. atm)

(Taken from the previous extreme chemisty grade 11 - 12 book)

Solution

$$CO_2 = K_H \cdot P_{CO_2}$$
$$= 3.3 \times 10^{-2} \text{ mol/L.atm} \times 4 \text{ atm}$$
$$= 0.132 \text{ mol/L}$$

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Illustrative Examples

32. In which type of solution does an equilibrium always exist?

A) Superaturated
B) Saturated
C) Unsaturated
D) Diluted

Solution

Dynamic equilibrium is established when the rate of dissolution is equal to the rate of crystallization. The solution that always exist at equilibrium is saturated solution.

Answer: B

33. A solution which has capacity to dissolve more solute is called

A) Unsaturated
B) Saturated
C) Supersaturated
D) Concentrated

Solution

Unsaturated solution dissolves additional solute at a given temperature.

Answer: A

34. When a small amount of crystal solute is added to the supersaturated solution, the solute crystal will:

A) Grow bigger
B) Slight dissolve
C) Remains unchanged
D) Dissolve completely

Solution

Supersaturated solution develops a crystal easily. As a result the size of the crystal increases.

Answer: A

35. Ambo water is prepared by injecting the liquid with carbon dioxide. Under what conditions is the solubility of carbon dioxide gas the greatest?

A) Low temperature, low pressure
B) Low temperature, high pressure
C) High temperature, high pressure

D) High temperature, low pressure

Solution

The solubility of a gas in liquid is the highest at low temperature and high pressure.

Answer: B

36. Which of the following is NOT correct?

A) Addition of a strongly solvated solute decreases the solubility of a gas in liquid
B) For solid dissolved in a liquid that is exothermic, increase in temperature increases the solubility.
C) The solubility of a substance is its concentration in the saturated solution.
D) For solution to occur, solvent - solute attractions must overcome solute solute and solvent - solvent attractions.

Answer: B

37. The heat of solution of sodium iodide (NaI) is -7500J/mol. At which temperature is NaI more soluble? 10°C or 40°C

Solution

Since $\Delta H_{sol} < 0$, NaI is more soluble at lower temperature.

Answer: 10°C

38. The solubility of potassium dichromate at 20°C and 40°C is 10g and 40g per 100g of water respectively. What is the sign of heat of solution (ΔH_{sol})?

Solution

ΔH_{sol} is positive because as temperature increases, solubility also increases.

39. The solubility of oxygen gas in water at 25°C and 1 atm pressure of oxygen is 0.041 g/L. The solubility of oxygen in water at 3atm and 25°C is ___ g/L.

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Solution

$$\frac{C_1}{C_2} = \frac{KH \cdot P_1}{KH \cdot P_2}$$

$$\Rightarrow \frac{C_1}{C_2} = \frac{P_1}{P_2}$$

$$C_2 = \frac{C_1 \times P_2}{P_1} = \frac{0.041\ g/L \times 3atm}{1\ atm}$$

$$C_2 = 0.123\ g/L$$

40. The solubility of sodium selenate. $Na_2 Se O_4$ is 84g/ 100g of water at 35°. The type of solution obtained by dissolving:

a) 92 g $Na_2 SeO_4$ in 200g H_2O is _____

b) 45 g $Na_2 SeO_4$ in 50g of H_2O is _____

Solution

84 g /100g = 42g/ 50g = saturated solution

a) 92 g /200g = 46 g/ 100g = unsaturated solution

b) 45g/ 50g = supersaturated solution

41. How can you prepare supersaturated solution?

Solution

Supersaturated solution is formed by saturating a solution at higher temperature.

42. What type of gases do not obey Henry's law?

Solution

Gases that form strong intermolecular forces such as hydrogen or chemically with water such as HCl, H_2S and NH_3 do not obey henry's law.

7.4 Ways of Expressing the Concentration of Solution

$$\text{Concentration} = \frac{\text{Amount of solute}}{\text{Amount of solution or solvent}}$$

☑ The concentration of a solution can be expressed qualitatively and quantitatively.

☑ **Qualitative (non - numerical) expression** - is expressing a solution simply by saying dilute or concentrated solution.

☑ **Quantitative (numerical) expression** - can be expressed by using five concentration units.

1. Mass percent
2. Mole fraction
3. Molarity
4. Normality
5. Molality

I. Percent by Mass or Volume

$$\text{Percent by mass of solute (\% W/W)} = \frac{m_{solute}}{m_{solute} + m_{solvent}} \times 100$$

$$\text{(\%W/W) solute} = \frac{m_{solute}}{m_{solution}} \times 100$$

m - mass

Percent by mass of solute + percent by mass of solvent = 100

* % - parts per hundred = $\frac{1}{100}$

* ppm - parts per million = $\frac{1}{10^6}$

* ppb - parts per billion = $\frac{1}{10^9}$

* $\text{ppm solute} = \frac{m_{solute}}{m_{solution}} \times 10^6 = \% \times 10^4$

➥ ppm solute $= \frac{m_{solute}}{m_{solution}} \times 10^9 = ppm \times 10^3 = \% \times 10^7$

➥ Percent by volume (% V/V) of solute $= \frac{\text{Volume of solute}}{\text{Volume of solution}} \times 100$

➥ Percent be mass/volume $= \frac{\text{Mass of solute}}{\text{Volume of solution}} \times 100$

Example

1. 20g of sugar is dissolved in 60g of water Calculate the concentration in
   a) Percent by mass b) ppm c) ppb

Solution

a) % by mass $= \frac{m_{sugar}}{m_{sugar} + m_{water}} \times 100$

$= \frac{20g}{(20+60)g} \times 100$

$= 25\%$

b) $ppm = \% \times 10^4 = 25 \times 10^4 = 2.5 \times 10^5$ ppm

c) $ppb = \% \times 10^7 = 25 \times 10^7 = 2.5 \times 10^8$ ppb

2. The label on a 0.5L bottle says 11% by volume. How many liters of alcohol does the wine contain?

Solution

Volume of alcohol $= 0.5L \times 11\%$

$= 0.5L \times \frac{11}{100}$

$= 0.055L$

2. Mole fraction (X)

☑ **Mole fraction** is the ratio of the moles of certain substance to the total moles in a mixture of substances.

☑ Solution that do not transimit electricity are called **non - electrolytes**

☑ Solutions that transmit electricity due to ions that carry charge are known as **electrolytes.**

i. Mole fraction for non - electrolytes

$$X_{solute} = \frac{n_{solute}}{n_{solute} + n_{solvent}}$$

$$X_{solvent} = \frac{n_{solvent}}{n_{solute} + n_{solvent}}$$

n - Number of moles

Note: sum of mole fraction of components of a mixture is one

➥ $X_{solute} + X_{solvent} = 1$

➥ Mole fraction is **unitless**

ii. Mole Fraction for Electrolytes

In electrolytic solutions, we have positive ion (cation), negative in (anion) and solvent. Thus,

$$x_{cation} = \frac{n_{cation}}{n_{cation} + n_{anion} + n_{solvent}}$$

$$x_{anion} = \frac{n_{anion}}{n_{cation} + n_{anion} + n_{solvent}}$$

$$x_{solvent} = \frac{n_{solvent}}{n_{cation} + n_{anion} + n_{solvent}}$$

$$\boxed{x_{cation} + x_{anion} + n_{solvent} = 1}$$

$$\boxed{\%\ mole = \frac{n_{solute}}{n_{solute} + n_{solvent}} \times 100}$$

Example

1. 2mol A and 3mol. B are mixed together to form a solution. Find the molecule fraction of A and B.

Solution

$$x_A = \frac{n_A}{n_A + n_B} = \frac{2mol}{(2+3)mol} = 0.4$$

$$x_B = \frac{n_B}{n_A + n_B} = \frac{3mol}{(2+3)} = 0.6$$

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or

$x_A + x_B = 1$

$x_B = 1 - x_A = 1 - 0.4 = 06$

2. 2mol $CaCl_2$ is dissolved in 4mol of water Calculate the mole fraction of Ca^{2+}, Cl^- and H_2O.

Solution

First, we have to find the number of moles of ions using stoichiometry.

$CaCl_2 \longrightarrow Ca^{2+} + 2Cl^-$

1mol 1mol 2mol

2mol ? ?

$$nCa^{2+} = \frac{2mol \times 1mol}{1mol} = 2mol$$

$$nCl^- = \frac{2mol \times 2mol}{1\ mol} = 4mol$$

$$\therefore XCa^{2+} = \frac{n_{ca}^{2+}}{nCa^{2+} + nCl^- + nH_2O} = \frac{2mol}{(2+4+4)\ mol} = 0.2$$

$$XCl^- = \frac{nCl-}{nCa^{2+} + nCl^- + nH_2O} = \frac{4mol}{(2+4+4)mol}$$

$XCl^- = 0.4$

$XH_2O = 1 - (xCa^{2+} + xCl^-)$

$= 1 - (0.2 + 0.4)$

$= 0.4$

Illustrative Examples

43. What is the concentration of a solution in parts per million, if 0.02g of Na_3PO_4 is dissolved in 1000g of water?

A) 0.02ppm B) 0.2ppm C) 2ppm D) 20ppm

Solution

$$ppm = \frac{m_{solute}}{m_{solute} + m_{solven}} \times 10^6$$

$$= \frac{0.02g \times 10^6}{(0.02 + 1000)g} = 20ppm$$

Answer: D

44. A beaker filled to the 100mol mark with salt (the mass of salt is 100g) and another beaker to the 100ml mark with water (the water has a mass of 100g) are mixed in a bigger beaker until the salt is completely dissolved. What will be the mass of the solution?

A) It will be much more than 200g

B) It will be much smaller than 200g

C) It will be exactly 200g

D) It will be slightly more than 200g

Solution

Mass of solution = mass of salt + mass of water

= 100g + 100g = 200g

Mass is additive but volume is not necessarily additive due to expansion and contraction.

Answer: C

45. A commercial bleaching solution ("berekina") contains mass percent of 5% sodium hypochlorite, NaOCl. What is the mass of NaOCl in a bottle containing 2500g of bleaching solution?

Solution

Mass of NaOCl = 5% × mass of solution

$$= \frac{5}{100} \times 2500g$$

$= 125g$

46. A 500g toothpaste sample has 0.2g fluoride concentration. What is the concentration of fluoride interms of ppm?

Solve by your own! Answer: 400ppm

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47. A certain vinegar is 6.02% acetic acid ($C_2H_4O_2$) by mass. How many grams of $C_2H_4O_2$ are contained in a 355ml bottle of vinegar. Assume a density of 1.01g/ml.

Solution

$$m = P \times V = 1.01\,g/ml \times 355ml = 358.55g$$

$$\%\,(w/w)\ acetic\ acid = \frac{mass\ of\ aceticacid}{mass\ of\ vinegar} \times 100$$

$$mass\ of\ acetic\ acid = \%\ x\ mass\ of\ vinegar$$

$$= \frac{6.02}{100} \times 358.55g$$

$$= 21.58g$$

48. A solution contains 0.8g of glucose in 100ml of solution. what is the % (W/V) of glucose of this solution?

Solution

$$\%\,(w/v) = \frac{m_{glucose}}{V_{solution}} \times 100$$

$$= \frac{0.8\,g}{100\,ml} \times 100$$

$$= 80\%$$

$$\boxed{100ml = 100g \quad or \quad 1ml = 1g}$$

49. If 150g of orange juice contains 120mg of ascorbic acid, what is the concentration of ascorbic acid expressed in

a) % by mass b) ppm c) ppb

Solution

a) $$\%\,(W/W)\ ascorbis\ acid = \frac{m_{gulcose}}{V_{solution}} \times 100$$

$$= \frac{0.12g}{150g} \times 100$$

$$= 0.08\%$$

b) $ppm = \% \times 10^4 = 0.08 \times 10^4 = 800ppm$

c) $ppb = \% \times 10^7 = 0.08 \times 10^7 = 8 \times 10^5$ ppb

50. What is the mole fraction of methanol in a mixture of 16g methanol (CH_3OH) and 27g of water (H_2O).

Solution

$$n_{methanol} = \frac{m_{methanol}}{M_{methanol}} = \frac{16g}{32g/mol} = 0.5mol$$

$$n_{water} = \frac{m_{water}}{M_{water}} = \frac{27g}{18g/mol} = 1.5mol$$

$$X_{methanol} = \frac{n_{methanol}}{n_{methanol} + n_{water}} = \frac{0.5mol}{(0.5 + 1.5)mol} = 0.25$$

$$X_{water} = 1 - X_{methanol} = 1 - 0.25 = 0.75$$

51. 3moles of sodium sulphide Na_2S, is dissolved in 1moles of water. Calculate the mole fraction of Na^+, S^{2-} and H_2O.

Solution

$$Na_2S \longrightarrow 2Na^+ + S^{2-}$$

1mol 2mol 1mol

3mol ? ?

$$nNa^+ = \frac{3mol \times 2mol}{1mol} = 6mol$$

$$nS^{2-} = \frac{3mol \times 2\ mol}{1mol} = 3mol$$

$$nH_2O = 1mol$$

$$XNa^+ = \frac{6mol}{(6 + 3 + 1)mol} = 0.6$$

$$XS^{2-} = \frac{3mol}{(6 + 3 + 1)mol} = 0.3$$

$$XH_2O = 1 - (0.6 + 0.3) = 0.1$$

52. Calculate the mole fractions of H_2O and CH_3OH in a mixture of equal masses of water and water methanol.

Solution

Let the mass be m

$$X_{methnol} = \frac{\frac{m}{M_{methanol}}}{\frac{m}{M_{methanol}} + \frac{m}{M_{water}}}$$ 1M-molar mass

$$= \frac{\frac{m}{32}}{\frac{m}{32} + \frac{m}{18}} = \frac{m}{32} \times \frac{32 \times 18}{18m + 32m}$$

$$X_{methanol} = \frac{18}{18+32} = \frac{18}{50} = \frac{9}{25}$$

$X_{methanol} = 0.36$

$X_{water} = 1 - 0.36 = 0.64$

For equal masses of two substance A and B, to determine their mole fraction use the simple formula below:

$$X_A = \frac{M_B}{M_A + M_B}$$ M_A - molar mass of A

$$X_B = \frac{M_A}{M_A + M_B}$$ M_B - molar mass of B

53. 5 moles of solute A is dissolved in 11moles of solvent B. Determine the mole percent of A and B.

Solution

$$\text{Mole \% } A = \frac{n_A}{n_A + n_B} \times 100 = \frac{5mol}{(5+11)mol} \times 100 = 31.25\%$$

$$Mole\ \%\ B = 100 - mole\ \%\ A = 68.75\%$$

Mole % $A = X_A \times 100$

Mole % $B = X_B \times 100$

3. Molarity (M)

☑ Molarity is the most common concentration unit in the laboratory.

☑ It is defined as number of mole of solute divided by volume of solution in liter or cubic decimeter.

$$\text{Molarity} = \frac{\text{Number of mole of solute}}{\text{Volume of solution in liter}}$$

$M = \frac{n}{V}$ $n = M \times V$

$n = \frac{m}{M_{wt}}$ $M = \frac{m}{M_{wt} \times V}$

m = given Mass of solute

M_{wt} = molar mass of solute

$1L = 1dm^3 = 1000ml = 1000cm^3$

☑ The unit of molarity is molar (M)

$1\,M = 1mol/L = 1mol/dm^3$

Example: What is the molarity of a solution obtained by dissolving 0.01mol of NaCl in 500ml of solution?

Solution

$n = 0.01mol$, $V = 500ml = 0.5L$

$$M = \frac{n}{V} = \frac{0.01mol}{0.5L} = 0.02mol/L$$

$M = 0.02M$

The precision of measuring molarity of solution in laboratory is affected by:

☑ Temperature - a solution expands when heated and contracted when cooled. As a result, the molarity differs.

☑ Mixing - volume may not be additive due to expansion or contraction.

Illustrative Examples

54. How many moles of H_2SO_4 are needed to prepare 5.0 liter of a solution of H_2SO_4 ?

A) 2.5 B) 5.0 C) 20 D) 10

Solution

$n = M \times V = 2mol/\cancel{L} \times 5\ \cancel{L} = 10\ mol$

Answer: D

55. What is the molarity of a 5g hydrogen peroxide (H_2O_2) in 100ml solution that is used for hair bleaching?

A) 0.015M B) 0.15M C) 1.5M D) 3M

Solution

$V = 100ml = 0.1L$

$M_{wt} = 34g/mol$

$m = 5g$

$$M = \frac{M}{M_{wt} \times V} = \frac{5g}{34\frac{g}{mol} \times 0.1L} = 1.5M$$

Answer: C

56. The concentrative of nitrate ion in a solution that contains 0.900M aluminum nitrate is:

A) 0.90 B) 0.45 C) 0.30 D) 2.70

Solution

The formula of aluminum nitrate is $Al(NO_3)$

$Al(NO_3)_3 \longrightarrow Al^{3+} + 3NO_3^-$

1M 3M

0.9M ?

$$[NO_3^-] = \frac{3\cancel{M} \times 0.900M}{1\cancel{M}} = 2.70M$$

Answer: D

57. What is the mass of $CuCl_2 . 6H_2O$ required to prepare 300ml of 0.10M aqueous solution?

Solution

$$m = M \times M_{wt} \times V = 0.10\frac{mol}{\cancel{L}} \times 242.5\frac{g}{mol} \times 0.3\cancel{L}$$

$m = 7.28g$

58. What volume of 5.0M sulphuric acid contains 0.1mol H_2SO_4 on milliliters?

Solution

$$V = \frac{n}{M} = \frac{0.1mol}{5\frac{mol}{L}} = 0.02L = 20ml$$

59. How many chloride ions are there in 1mole of calcium chloride, $CaCl_2$?

Solution

$n_{Cl^-} = 2 \times nCaCl_2 = 2 \times 1mol = 2mol$

$CaCl_2 \longrightarrow Ca^{2+} + 2Cl^-$

1mol 2mol

$\Rightarrow n_{Cl^-} = 2mol$

$$\text{Number of } Cl^- \text{ ions} = nCl^- \times 6.02 \times 10^{23}\ \frac{ions}{mol}$$

$$= 2mol \times 6.02 \times 10^{23}\ \frac{ions}{mol}$$

$$= 1.204 \times 10^{24}\ ions$$

4. Normality (N)

Normality is defined as the number of equivalents of solule per liter of solution.

$$\text{Normality} = \frac{\text{Number of equivalents of solute}}{\text{Volume of solution in liter}}$$

$$\boxed{N = \frac{No.\ eq}{V}}$$

$$No\ eq = \frac{m}{E}$$

m - given mass

E - Equivalent mass

$$E = \frac{M_{wt}}{Z}$$

M_{wt} - molar mass

Z - number of hydrogen ion for acid
- ➥ Number of hydroxide ion for base
- ➥ Total positive charge for salt
- ➥ Electrons lost for reducing agent
- ➥ Electron gained for oxidizing agent.

$$No.\ eq = \frac{m}{E} = \frac{m \times Z}{M_{wt}}$$

$No.\ eq = n \times Z$ because $n = \frac{m}{M_{wt}}$

$$N = \frac{m}{E.V} = \frac{m \times Z}{M_{wt} \times V} = \frac{n \times Z}{V}$$

$N = M \times Z$ because $M = \frac{n}{V}$

☑ The unit of normality is equivalent per liter (eq/L) or normal (N)

1N = 1eq//L

☑ The unit of equivalent mass is gram per equivalent (g/eq).

Normality depends on the type of reaction due to the value of Z

Example

1. Determine the equivalent mass of
   a) H_2SO_4 b) $Ca(OH)_2$ c) Na_2S

Solution

a) $H_2SO_4 \longrightarrow 2H^+ + SO_4^{2-}$

$Z = \frac{2eq}{mol}$

$MH_2SO_4 = 98g/mol$

$E = \frac{M_{wt}}{Z} = \frac{98\ g/mol}{2eq/mol} = 49\ g/eq$

b) $Ca(OH)_2 \longrightarrow Ca^{2+} + 2OH^-$

$Z = 2eq/mol$

$MCa(OH)_2 = 74g/mol$

$E = \frac{M_{wt}}{Z} = \frac{74g/mol}{2eq/mol} = 37g/eq$

c) $Na_2S \longrightarrow 2Na^+ + S^{2-}$

$Z = 2 \times (1) = 2\frac{eq}{mol}$

$MNa_2S = 78g/mol$

$E = \frac{M_{wt}}{Z} = \frac{78g/mol}{2eq/mol} = 39g/eq$

2. Determine the normality of a solution prepared by dissolving 16g of $Fe_2(SO_4)_3$ in 200ml of solution.

Solution

$Fe_2(SO_4)_3 \longrightarrow 2Fe^{3+} + 3SO_4^{2-}$

$Z = 2 \times (3) = 6eq/mol$

$M_{wt} = 400g/mol$

$V = 200ml = 0.2L$

$$N = \frac{m \times Z}{M_{wt} \times V} = \frac{16g \times 6eq/mol}{400\frac{g}{mol} \times 0.2L}$$

$N = 1.2N$

5. Molality (m)

The molality of a solution is the number of mole of solute per mass of solvent in kilogram

$$m = \frac{n}{m_{solvent}\ in\ kg}$$

The unit of molality is mole per kilogram or molal

$$\boxed{1m = 1\frac{mol}{kg}}$$

To prepare molal solution, we need to measure the mass of solvent not the amount of solution.

Illustrative Example

60. What is the normality of a 2.3M H_2SO_4 solution?

A) 1.15N B) 2.3N C) 4.6N D) 6.9N

Solution

$$H_2SO_4 \longrightarrow 2H^+ + SO_4^{2-}$$

$$Z = \frac{2eq}{mol}$$

$$N = M \times Z = 2.3\frac{møl}{L} \times \frac{2eq}{møl}$$

$$N = 4.6N$$

Answer: C

61. The concentration unit that depends on the reaction is

A) Molarity C) Molality

B) Normality D) Mole fraction

Answer: B

62. For which of the following concentration unit, it is not necessary to measure the quantity of solution?

A) Mass percent C) Molarity

B) Normality D) Molality

Answer: D

63. Calculate the number of equivalent present in 1.5mol H_3A if the acid is

a) Completely ionized to give A^{3-}

b) Converted to HA^{2-}

c) Converted to H_2A^-

Solution

a) $H_3A \longrightarrow 3H^+ + A^{3-}$

$$Z = 3\frac{eq}{mol}$$

$$\text{No eq} = n \times Z = 1.5møl \times \frac{eq}{møl} = 4.5eq$$

b) $H_3A \longrightarrow HA^{2-} + 2H^+$

$$Z = 2\frac{eq}{mol}$$

$$\text{No eq} = n \times Z = 1.5møl \times 2\frac{eq}{møl} = 3.0eq$$

c) $H_3A \longrightarrow H_2A^- + H^+$

$$Z = 1eq/mol$$

$$\text{No eq} = n \times Z = 1.5møl \times \frac{eq}{mol} = 1.5eq$$

64. Determine the equivalent mass of H_3PO_4 based on the reactions below.

a) $H_3PO_4 + NaOH \rightarrow Na\,H_2PO_4 + H_2O$

b) $H_3PO_4 + 2NaOH \rightarrow Na_2HPO_4 + 2H_2O$

c) $H_3PO_4 + 3NaOH \rightarrow Na_3PO_4 + 3H_2O$

Solution

a) Only one hydrogen is replaced with Na

$\Rightarrow Z = 1eq/mol$

$$E(H_3PO_4) = \frac{M_{wt}}{Z} = \frac{98g/mol}{1eq/mol} = 98g/eq$$

b) Two the hydrogen atoms are replaced with two Na atoms

$\Rightarrow Z = 2eq/mol$

$$E(H_3PO_4) = \frac{M_{wt}}{Z} = \frac{98g/mol}{2eq/mol} = 32.7g/eq$$

c) Three of the hydrogen are replaced with Na

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$\Rightarrow Z = 3 eq/mol$

$$E(H_3PO_4) = \frac{M_{wt}}{Z} = \frac{98g/mol}{3eq/mol} = 34.7g/eq$$

65. What volume of a 0.35N solution contains 4m.eq of solution?

Solution

$$N = \frac{No.eq}{V}$$

$$V = \frac{No.eq}{N} = \frac{4m\,eq}{0.35\frac{eq}{L}} = 11.4mL$$

66. Calculate the mass of $CuSO_4$ in 250 ml of solution if concentration is 0.65N.

Solution

$$N = \frac{m \times Z}{M_{wt} \times V}, \quad Cu\,SO_4 \longrightarrow Cu^{2+} + SO_4^{2-}$$

$$Z = 2eq/mol$$

$$m = \frac{N \times M_{wt} \times V}{Z} = \frac{0.65\frac{eq}{L} \times 160\frac{g}{mol} \times 0.25L}{2eq/mol} = 13g$$

67. How many equivalents of solute is contained in 2.5L of 0.8N solution?

Solution

$$N = \frac{No.\,eq}{V}$$

$$No.eq = N \times V = 0.8\frac{eq}{L} \times 2.5L$$

$$No.eq = 2.00eq$$

68. What is the molality a solution obtained by dissolving 11.5g ethanol in 400g of water.

Solution

$$mH_2O = 400g = 0.40kg$$

$$n_{ethanol} = \frac{m}{M_{wt}} = \frac{11.5g}{46g/mol} = 0.25mol$$

$$m = \frac{n}{mH_2O} = \frac{0.25mol}{0.40kg} = 0.625m$$

69. Calculate the mass of $MgCl_2$ required to prepare 2.5m $MgCl_2$ solution in 200g of water.

Solution

$$m = \frac{m(MgCl_2)}{M_{wt} \times kg\ of\ H_2O}$$

$$m(MgCl_2) = m \times M_{wt} \times kg\ of\ H_2O$$

$$= 2.5\frac{mol}{kg} \times \frac{95g}{mol} \times 0.2kg$$

$$= 47.5g$$

70. What mass of water is required to dissolve 14.5g of glucose, $C_6H_{12}O_6$ to get a solution with a concentration of 0.24m?

Solution

$$m = \frac{m(glucose)}{M_{wt} \times kg\ H_2O}$$

$$kg\ H_2O = \frac{m(glucose)}{M_{wt} \times m}$$

$$= \frac{14.5g}{180\frac{g}{mol} \times 0.24\frac{mol}{kg}} = 0.336kg = 336g$$

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7.5 Conversion of Concentration Units

It is possible to convert one concentration unit into another based on their definitions. For example

- Mass percent and density into molarity, normality and molality.
- Mass percent into molality and mole fraction.
- Mole fraction into molality.
- Molality into mole fraction.
- Molarity and density into molality.

Illustrative Examples

71. What is the molarity of potassium iodide in a solution that is 5% by mass KI and has a density of 1.038g/ml?

A) 0313M B) 0.626M C) 0.500M D) 0.123M

Solution

We may used two methods to solve the problem. **method I**

The volume of solution is 1L or 1000mL from definition.

$$\Rightarrow m_{solution} = \rho \times V = 1.038\frac{g}{ml} \times 1000ml = 1038g$$

$$m_{KI} = 5\% \times m_{solution} = \frac{5}{100} \times 1038g = 51.9g$$

$$M = \frac{m_{KI}}{M_{KI \times V}} = \frac{51.9g}{166\frac{g}{mol} \times 1L} = 0.313M$$

Method II

$$M = \frac{\% \times \rho \times 10}{M_{Solute}} = \frac{5 \times 1.038 \times 10}{166} = 0.313M$$

Answer: A

72. What is the molality of a solution that contains 51.2g of naphthalene, $C_{10}H_{18}$ in 500ml of CCl_4 ? $\left(\text{Density of } CCl_4 = 1.60\frac{g}{ml}\right)$

A) 0.750m B) 0.500m C) 0.840m D) 1.69m

Solution

$$mCCl_4 = \rho \times V = 1.60\frac{g}{ml} \times 500ml = 800g$$

$$= 0.80kg$$

$$m = \frac{m(C_{10}H_{18})}{MC_{10}H_{18} \times kg\ CCl_4} = \frac{51.2g}{128\frac{g}{mol} \times 0.8kg} = 0.50m$$

Answer: B

73. An aqueous solution is 70% nitric acid (HNO_3) by mass. What is the concentration of HNO_3, expressed in molality?

A) 0.55m B) 8.62m C) 11.1m D) 37m

Solution

70% by HNO_3 means 70g HNO_3 is dissolved in 30g of H_2O

$$\Rightarrow m = \frac{mHNO_3}{MHNO_3 \times kgH_2O} = \frac{70g}{63\frac{g}{mol} \times 0.030kg} = 37.04m$$

Answer: D

74. What is the mole fraction of a solute in one molal aqueous solution?

A) 0.009 B) 0.018 C) 0.027 D) 0.036

Solution

$$1m = 1\frac{mol}{kg} \longrightarrow \text{definition}$$

$$\Rightarrow n_{solute} = 1mol$$

$$m(H_2O) = 1kg = 1000g$$

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$$n(H_2O) = \frac{m}{M} = \frac{1000g}{18g/mol} = 55.6mol$$

$$X_{solute} = \frac{n_{solute}}{n_{solute} + nH_2O}$$

$$= \frac{1mol}{(1 + 55.6)\ mol} = 0.0176$$

Answer: B

75. Calculate the molality of a solution of ethanol in water. Mole fraction of ethanol is 0.25.

Solution

Consider the mole fraction as number of mole

$X_{ethanol} = n_{ethanol} = 0.25mol$

$XH_2O = 1 - 0.25 = 0.75 \Rightarrow nH_2O = 0.75mol$

$$\Rightarrow mH_2O = n \times M_{wt} = 0.75mol \times 18\frac{g}{mol} = 13.5g$$

$$= 0.0135kg$$

$$\therefore m = \frac{n}{kgH_2O} = \frac{0.25mol}{0.0135kg} = 18.5m$$

76. A solution contains equal masses of glycerol $C_3H_8O_3$, and water. The density of the solution is 1.10g/ml. What is the molality of the solution?

Solution

$V_{solution} = 1000ml$

$$m_{solution} = \rho \times V = \frac{1.10g}{ml} \times 1000ml = 1100g$$

$$m_{water} = m_{grycerol} = \frac{1100g}{2} = 550g$$

$$m = \frac{m_{gly}}{m_{gly} \times kg\,H_2O} = \frac{550g}{92\frac{g}{mol} \times 0.55kg} = 10.90m$$

77. A 2.45 M sulphuric acid solution has a density of 1.15 g/ml at 20°C.

calculate.

a) Normality

b) Molality

c) Percentage by mass of H_2SO_4

d) Mole fraction of H_2SO_4

Solution

a) $$N = M \times Z = 2.45\frac{mol}{L} \times 2\frac{eq}{mol} = 4.90N$$

b) $$2.45M = 2.45\frac{mol}{L}$$

$n = 2.45mol$

$V = 1L = 1000ml$

$$m_{solution} = \rho \times V = 1.15\frac{g}{ml} \times 1000ml = 1150g$$

$$m(H_2SO_4) = n \times Mwt = 2.45mol \times 98\frac{g}{mol} = 240g$$

$mH_2O = m_{solution} - mH_2SO_4$

$= 1150g - 240g = 910g = 0.910kg$

$$m = \frac{n}{kgH_2O} = \frac{2.45mol}{0.910kg} = 2.69m$$

c) $$\%\,H_2SO_4 = \frac{mH_2SO_4}{m_{solution}} \times 100$$

$$= \frac{240g}{1150g} \times 100 = 21\%$$

d) $$nH_2O = \frac{m}{m} = \frac{910g}{18g/mol} = 50.60mol$$

$$X(H_2SO_4) = \frac{nH_2SO_4}{nH_2SO_4 + nH_2O} = \frac{2.45mol}{(2.45 + 50.6)mol} = 0.046$$

78. What is the molarity of a300ml solution which is 25% by mass H_2SO_4? The solution mass is 353.4g.

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Solution

$$MH_2SO_4 = 25\% \times M_{solution} = \frac{25}{100} \times 353.4g = 88.35g$$

$$M = \frac{mH_2O_4}{MH_2O_4 \times V} = \frac{88.35g}{98\frac{g}{mol} \times 0.3L} = 3.0M$$

79. A solution of lead (II) nitrate, $Pb(NO_3)_2$, contains 18% $Pb(NO_3)_2$ by mass. If the density of the solution is 1.18g/cm³. Find

a) Molarity b) Molality c) Normality

Solution

a) $$M = \frac{\rho \times \% \times 10}{M_{solute}} = \frac{1.18 \times 18 \times 10}{331} = 0.64M$$

b) Use 18% by mass $Pb(NO_3)_2$ and solve molality

$m_{solute} = 18g$

$\Rightarrow$ 82g H_2O = 0.082kg

$$m = \frac{18g}{331\frac{g}{mol} \times 0.082kg} = 0.66m$$

c) $N = M \times Z = 0.64 \times 2 = 1.28N$

80. 10% aqueous solution of certain substance is isotonic with 5% aqueous solution of fructose ($C_6H_{12}O_6$). What is the molecular mass of the substance?

Solution

Isotonic solutions are solutions that have the same concentration. For two solutes A and B for isotonic solutions;

$$\frac{\%A}{M_A} = \frac{\%B}{M_B}$$

$$\frac{10\%}{M_{sub}} = \frac{5\%}{M_{fructose}}$$

$$\frac{10}{M_{sub}} = \frac{5}{180g/mol}$$

$M_{sub} = 360g/mol$

7.6 Solution Preparation

☑ A solution can be prepared in schools, hospitals and other areas.

☑ To prepare a solution from solid solutes follow the steps below:

1. Weigh the solute accurately
2. Transfer the weighed solute into volumetric flask
3. Add water using funnel
4. Dissolve the solid gently.
5. Add more water to bring the level of solution exactly to the volumetric mark.
6. Lable your Solution

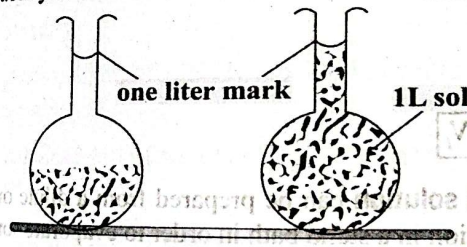

Fig 7.2 Preparing a Molar Solution

Example : How could you prepare 200cm³ of 0.5M of NaOH solution?

Solution

First, determine the mass of the solute (NaOH)

$m = M \times MNaOH \times V$

$$= 0.5\frac{mol}{L} \times 40\frac{g}{mol} \times 0.2L = 4g$$

Thus, transfer 4g of NaOH in a 200cm³ volumetric flask, dissolve the solute and add enough water upto the maker carefully.

The type of water we use to prepare a solution should be distilled Water

Dilution and Concentration

☑ **A concentrated solution** contains more solute where as, a dilute solution contains fewer solute in a given volume of solution.

☑ To prepare a dilute solution from a concentrated solution, a given volume of concentrated solution is measured and more solvent is added.

During Dilution

➥ Mole of solute is unaffected

➥ Volume and molarity change

Mathematically

$$n_i = n_f$$
$$\Rightarrow M_i V_i = M_f V_f$$

i - initial

f - final

$$V_{\text{solvent added}} = V_f - V_i$$

☑ **A more concentrated solution** can be prepared from a dilute one by heating the dilute solution in a stand bath in order to evaporate some of the solvent.

$$V_{\text{solvent evaporated}} = V_i - V_f$$

To diute concentrated acid obey the laboratory safety rule, add acid to water

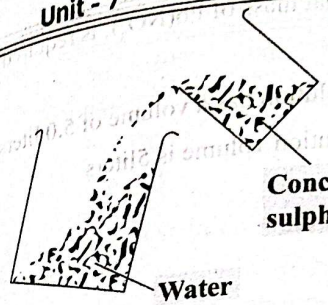

Fig 7.3 Adding concentrated sulphuric acid to water

Illustrative Examples

81. If a student wishes to prepare appropriately 100ml of an a aqueous solution of 6M HCl using 12HCl, which procedure is correct?

A) Adding 50ml of 12M HCl to 50ml of water while stirring the mixture steadily.

B) Adding 25ml of 12N HCl to 50ml of water while stirring the mixture steadily.

C) Adding 50ml of water to 50ml of 12MHCl while stirring the mixture steadily.

D) Adding 25ml of water to 50ml of 12MHCl while stirrng the mixture steadily.

Solution

$M_i = 12M$ $\quad M_f = 6M$

$V_i = ?$ $\quad V_f = 100ml$

$M_i V_i = M_f V_f$

$$V_i = \frac{M_f V_f}{M_i} = \frac{6M \times 100ml}{12m} = 50ml$$

$VH_2O = V_f - V_i = 100ml - 50ml = 50ml$

Add acid to water not water to acid! **Answer: A**

82. A laboratory instructor is preparing 5.0 liters of a 0.10M $Pb(NO_3)_2$

(molecular mass 331) Solution. What mass of $Pb(NO_3)_2$ is required?

A) 165.5g and add 5kg of H_2O

B) 165.5g and add H_2O until the solution has a volume of 5.0 liters.

C) 33.1g and add H_2O until the solution volume is 5liters

D) 33.1g and add 5L of water.

Solution

$$m = M \times M_{wt} \times V$$

$$= 0.1 \frac{mol}{L} \times 331 \frac{g}{mol} \times 5L = 165.5g \text{ of } Pb(NO_3)_2$$

Answer: B

83. How many ml of water is required to dilute 50ml of 3.5M H_2SO_4 solution to 2M?

Solution

$$V_f = \frac{M_i V_i}{M_f} = \frac{3.5M \times 50ml}{2M} = 87.5ml$$

$$VH_2O \text{ added} = V_f - V_i = 87.5ml - 50ml = 37.5ml$$

84. How much water has to be evaporated from 250ml of 1M to make it 3M?

Solution

$$V_f = \frac{M_i V_i}{M_f} = \frac{1M \times 250ml}{3M} = 83.3ml$$

$$VH_2O \text{ evaporated} = V_i - V_f = (250 - 83.3)ml$$

$$= 166.7ml$$

85. How could you prepare a 5% by mass KI solution in 40g of water?

Solution

Let the mass of KI be m

$$\% KI = \frac{m}{m + mH_2O}$$

$$\frac{5}{100} = \frac{m}{m + 40}$$

$$5m + 200 = 100m$$

$$95m = 200$$

$$m = \frac{200}{65} = 2.1g$$

$\therefore$ 5% by mass KI solution is prepared by dissolving 2.1g of KI in 40g of water.

86. Calculate the final concentration of a solute when 2L of 3M sugar solution and 3L of 2.5M sugar solution are mixed and then diluted to 10L with water.

Solution

Solution 1: $n_1 = M_1 V_1 = 3M \times 2L = 6mol$

Solution 2: $n_2 = M_2 V_2 = 2.5M \times 3L = 7.5mol$

$$n_T = n_1 + n_2 = 6 + 7.5 = 13.5mol$$

$$M_f = \frac{n_T}{V_T} = \frac{13.5mol}{10L} = 1.35 \text{ M}$$

87. What is the volume of water that must be added to a mixture of 200ml of 5M HCL and 800ml of 3M to obtain a 2M solution?

Solution

$$M_1 V_1 + M_2 V_2 = M_3 V_3$$

$$V_3 = \frac{M_1 V_1 + M_2 V_2}{M_3} = \frac{200 \times 5 + 800 \times 3}{2} = 1700ml$$

$$V_3 = 1700ml$$

$$VH_2O \text{ added} = V_3 - (V_1 + V_2) = 1700 - (800 + 200)$$

$$= 700ml$$

88. How do you prepare 30ml of 4M HNO_3 from a stock solution of 16M HNO_3?

Solution

$$V_i = \frac{M_f\ V_f}{M_i} = \frac{4M \times 30ml}{16m} = 7.5ml$$

➥ You should dilute 7.5ml of 16MHNO$_3$ with sufficient water to give the final volume 30ml.

7.7 Solution Stoichiometry

Stoichiometry is the study of the quantitative relationship between amounts of reactants and reactants, reactants and products, or products and products in a balanced chemical equation.

Solution stoichiometry differs from pure substance stoichiometry in that, the number of mole of pure substance is obtained when given mass is divided by molar mass. However, the number of mole of solute is the product of molarity and volume.

For pure substance:

$$n = \frac{m}{M_{wt}},$$

m - given mass

M_{wt} molar mass

For solution

$n = M \times V$, M - molarity

V - volume

There are three types of solution stoichiometry

1. Mole - Mass Relationship

From a given mole or volume and molarity, it is possible to determine mass.

Example

How many grams of zinc react with 500ml of 1M HCl to form zinc chloride and hydrogen gas?

Solution

➥ First, write the balanced chemical equation

$Zn(s) + 2HCl\ (aq) \rightarrow ZnCl_2\ (aq) + H_2(g)$

➥ Second, determine the number of mole of HCl

$$n_{HCl} = M \times V = 1\ \frac{mol}{L} \times 0.5L = 0.5mol$$

➥ Third, by mole - mass relationship, find the mass of Zn.

$$\overset{x}{Zn}\ (s) + \overset{0.5mol}{2HCl}\ (aq) \longrightarrow ZnCl_2\ (aq) + H_2\ (g)$$

$1 \times 65g \qquad 2mol$

$$\frac{x}{65g} = \frac{0.5mol}{2mol},\ x = \text{mass of } Zn = 16.25g$$

Thus, the mass of zinc that reacts with 500ml of 1M HCl is 16.25g

2. Mole - Volume Relationship

From a give molarity and volume, we can find number of mole or vice versal.

Example

What volume of 0.5M KOH solution is needed to react with 50ml of 0.5M H_2SO_4 solution?

Solution

➥ Write balanced chemical equation

$2KOH\ (aq) + H_2SO_4\ (aq) \rightarrow K_2SO_4\ (aq) + 2H_2O\ (1)$

➥ Find number of mole of H_2SO_4

$$n = M \times V = 0.5\ \frac{mol}{L} \times 0.05L = 0.025mol$$

➥ By mole - mole relationship find number of mole of KOH

$$\overset{x}{2KOH}\ (aq) + \underset{1}{H_2SO_4}\ (aq) \rightarrow K_2SO_4\ (aq) + 2H_2O(1)$$

$$\frac{x}{2} = \frac{0.025mol}{1}$$

$x = 0.05mol = n_{KOH}$

➥ From mole and molarity, find volume of KOH.

$$V_{KOH} = \frac{n}{M} = \frac{0.05mol}{0.5mol/L} = 0.1L = 100ml$$

Thus, the volume of 0.5M KOH solution that reacts completely with 50ml of 0.5M of H_2SO_4 solution is 100ml.

3. Mole - Number of Particles Relationship

A mole is Avogadro's number of particles.

1mole = 6.02×10^{23} particles (atoms molecules ions formula units)

Example

What volume of 0.125M sucrose ($C_{12}H_{22}O_{11}$) contains 1.204×10^{23} molecules of sucrose?

Solution

1mole = 6.02×10^{23} molecules

? = 1.204×10^{23} molcules

$$n = \frac{1mol \times 1.204 \times 10^{23} \text{ molecules}}{6.02 \times 1023 \text{ molecules}}$$

$$V = \frac{n}{M} = \frac{0.2mol}{0.125\frac{mol}{L}} = 1.6L$$

Illustrative Example

89. What volume of 0.50M NaOH is required to neutralize 25ml of 1.2M H_2SO_4 solution? (Assume complete ionization of the acid)

A) 60ml B) 90ml C) 100ml D) 120ml

Solution

$$n(H_2SO_4) = M \times V = 0.025L \times 1.2\frac{mol}{L} = 0.030mol$$

0.03mol x

$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$

1 2mol

$$\frac{0.030mol}{1mol} = \frac{x}{2mol}$$

$x = 0.060mol = n(NaOH)$

$$V_{NaOH} = \frac{n_{NaOH}}{M_{NaOH}} = \frac{0.06mol}{0.5\frac{mol}{L}} = 0.12L = 120ml$$

Answer: D

Other method

One equivalent of acid alwater reacts with

One equivalent of base

$\Rightarrow \boxed{N_a V_a = N_b V_b}$

a - acid

b - base

$$N_a = M \times Z = 1.2\frac{mol}{L} \times \frac{2eq}{mol} = 2.4N$$

$$N_b = M_b \times Z = 0.5\frac{mol}{L} \times 1\frac{eq}{mol} = 0.5N$$

$N_a V_a = N_b V_b$

$$V_b = \frac{N_a V_a}{N_b} = \frac{2.4N \times 25ml}{0.5N} = 120ml$$

$\boxed{M_a V_a \neq M_b V_b}$

90. How many grams of $NaHCO_3$ (baking soda) would you need to neutralize 500ml of battery acid (H_2SO_4) that has spilled on your floor? Assume that the concentration of the battery acid is 12molar. The balance chemical equation is:

$H_2SO_4(aq) + 2NaHCO_3(aq) \rightarrow Na_2SO_4(aq) + 2CO_2(g) + 2H_2O(l)$

A) 1.008×10^3kg C) 0.504kg

B) 1.008kg D) 0.252kg

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Solution

$$n(H_2SO_4) = M \times V = 12\frac{mol}{2} \times 0.5L = 6mol$$

6mol x

$$H_2SO_4(aq) + 2NaHCO_3(aq) \longrightarrow$$

1mol 2mol

$$\frac{6mol}{1mol} = \frac{x}{2mol}$$

$$x = n(NaHCO_3) = 12mol$$

$$m(NaHCO_3) = n \times M_{wt} = 12mol \times 84g/mol$$

$$= 1008g = 1.008kg$$

Answer: B

91. How many grams of $CaCO_3$ will be formed by passing excess CO_2 gas through 100ml of 0.25M $Ca(OH)_2$ solution?

$$Ca(OH)_2(aq) + CO_2(s) \rightarrow CaCO_3(s) + H_2O(l)$$

Solve by your own! Answer: 2.5g

92. How many HI particles are there in 0.2mol of HI solution?

Solution

1mol HI = 6.02×10^{23} HI molecules

0.2mol HI = ?

$$\text{No of HI molecules} = \frac{0.2mol \times 6.02 \times 10^{23}}{1mol}$$

$$= 1.204 \times 10^{23} \text{ HI particles}$$

7.8 Describing Reactions in Solution

Chemists use three types of equations to represent aqueous ionic equations. These are:

1. Molecular equation
2. Total ionic equation
3. Net ionic equation

1. Molecular equation shows species that are actually in solution as they

ions are present.

Example

$$NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

2. **Total ionic equation** reveals all the soluble ionic substances as they actually in solution where there are dissociated into ions. For the above example, the total ionize equation becomes:

$$Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$$

Note: pure solids, liquids and gases do not ionize and write them in molecular form.

3. **The net ionic equation** is very important because it eliminates the spectator ions and shows only the actual chemical change.

☑ **Spectator ions** are ions present in the total ionic equation but not in the net ionic equation

☑ They are not involved in actual chemical change but are only as part of the reactants or products.

☑ **$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$** is the net ionic equation for the above example

☑ Na^+ and NO_3^- are spectator ions.

Illustrative Examples

93. What is the balanced net ionic equation for the reaction of $CaCl_2$ and $AgNO_3$?

Solution

$$CaCl_2(aq) + 2AgNO_3(aq) \rightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$$

Molecular equation

$$Ca^{2+}(aq) + 2Cl^-(aq) + 2Ag^+(aq) \rightarrow 2AgCl(s) + Ca^{2+}(aq) + 2NO_3^-(aq)$$

Total ionic equation

$$2Cl^-(aq) + 2Ag^+(aq) \rightarrow 2AgCl(s)$$, net ionic equation

94. Consider the following balanced equation:

$$3Ba(NO_3)_2(aq) + Fe_2(SO_4)_3(aq) \rightarrow 2Fe_2(NO_3)_3(aq) + 3BaSO_4(s)$$

The net ionic equation to describe this balanced equation is _____

Answer: $3Ba^{2+}(aq) + 3SO^{2-}_4(aq) \rightarrow 3BaSO_4(s)$

95. Write the net ionic equation by identifying the spectator ion(s) for the following equations.

a) $2HNO_3(aq) + Ca(HCO_3)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2HO(1) + 2CO_2(g)$

b) $5FeCl_2(aq) + KMnO_4(aq) + 8HCl(aq) \rightarrow 5FeCl_3(aq) + MnCl_2(aq) + KCl(aq) + 4H_2O(1)$

c) $La_2(CO_2)_3(s) + 6HCl(aq) \rightarrow 2LaCl_3(aq) + 2CO_2(s) + 3H_2O(1)$

Solution

a) $2H^+(aq) + HCO^-_3(aq) \rightarrow 2H_2O(1) + 2CO_2(s)$

Ca^{2+} and NO^-_3 are spectator ions

b) $5Fe^{2+}(aq) + MnO^-_4(aq) + 8H^-(aq) \rightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(1)$

Spectator ion = Cl^-

c) $La_2(CO_3)_3(s) + 6H^+(aq) \rightarrow 2CO_2(g) + 3H_2O(1) + La^{3+}(aq)$

Spectator ion = Cl^-

Solved Problems

96. Explain the following

a) Isotonic solution
b) Hypertonic solution
c) Hypotonic solution
d) Coagulation
e) Association colloids

Solution

a) It is a solution that has same concentration

b) It is a solution that has a higher solute concentration

c) It is a solution that has lower solute concentration compared to another solute

d) Coagulation is the process by which a liquid changes from a fluid to a gel or thicker liquid.

➥ It is important for stopping blood loss from a damage vessel, forming a blood clot.

e) Association colloids are micro - heterogeneous particles of the colloidal dispersed phase (mi celles) are formed by asglomeration of molecules or ions of the substance assolved in the dispersed

medium.

97. The Henry's law constant of gas A is 0.12 mol/L.atm at 25°C. What is the solubility of A in water at 25°C at a partial pressure of 0.20atm?

Solution

$$C_A = K_H \cdot P_A = 0.12 \frac{mol}{L.atm} \times 0.20 atm = 0.024 mol$$

98. When you open Ambo water or soft drink, you hear a hissing sound and also observe bubbles escaping out of the water or drink. What does the escape of the gas bubbles indicate?

Solution

When we open the bottle containing ambo water or soft drink, the partial pressure of the gas decreases. As a result solubility of a gas in water decreases, as partial pressure of the gas decreases.

99. For an ideal solution the value of heat of solution (Δ Hsol) ___

Answer: $\Delta H_{sol} = 0$

100. Why does the soubility of any gas in water reduce with increasing temperature?

Solution

Because gases mix with water exothermically.

101. Classify the following substances as strong electrolytes, weak electrolytes or non - electrolytes.

a) Sodium permanganet $NaMnO_4$
b) Glucose ($C_6H_{12}O_6$)
c) Hydrochloric acid (HCl)
d) Ammonia (NH_3)
e) Sodium hydroxide (NaOH)
f) Methanol (CH_3OH)
g) Acetic acid (CH_3COOH)

Solution

☑ Strong acids, strong bases, molten and soluble salts are strong

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electrolytes.

☑ Weak acids, weak bases and partially soluble salts are weak electrolytes. other substances are non - electrolytes

- a, c and e are strong electrolytes
- d and g are weak electrolytes
- b and f are non - electrolytes.

102. Cow's milk typically contains 4.5% by mass of lactose $C_{12}H_{22}O_{11}$. Calculate the mass of lactose present in 175g of milk.

Solution

Mass of lactose = 4.5% × mass of milk

$$= \frac{4.5}{100} \times 175g = 7.875g$$

103. Write concentration units that are affected by temperature.

Solution

Molarity, normality and mass percent.

104. What is the molarity of sulphuric acid solution with a density of 1.84 g/cc at 35°C and contains 98% by weight sulphic acid?

Solution

$$M = \frac{\rho \times \% \times 10}{M_{wt}} = \frac{1.84 \times 98 \times 10}{98} = 18.4M$$

105. Two beakers contain 1L and 200mL of pure water respectively. What is the same in them?

Solution

The molarity of pure water is the same in both beakers, it is 55.6M.

106. A sample of sea water was found to contain 5×10^{-3}g of dissolved oxygen perkilogram. What is the concentration of oxygen in ppm?

Solution

$$\text{ppm } O_2 = \frac{\text{Mass of } O_2}{\text{Mass of seawater}} \times 10^6 = \frac{5 \times 10^{-3} g}{1000g} \times 10^6 = 5ppm$$

107. 10ml of 0.1M HCl is mixed with 5ml of 0.2M HNO_3 and 20ml of 0.3M H_2SO_4. What is the normality of the acids mixture

Solution

$$n(HCl) = M \times v = \frac{0.mol}{\not{L}} \times = 1mmol$$

$$\text{No.eq (HC)} = n \times Z = 1mmol \times \frac{1eq}{mol} = 1meq$$

$$n(HNO_3) = M \times V = 5mL \times 0.2\frac{mol}{L} = 1mmol$$

$$\text{No . eq } (HNO_3) = n \times Z = 1mmol \times \frac{1eq}{mol} = 1meq$$

$$n(H_2SO_4) = M \times V = 20mL \times 0.3\frac{mol}{L} = 6mmol$$

$$\text{No. eq } (H_2SO_4) = n \times Z = 6mmol \times 2\frac{eq}{mol} = 12meq$$

mmol = milli mole, meq = millii equivalent

(No.eq)total = 1meq + 1meq + 12meq = 14meq

$$N_T = \frac{(\text{No-eq})\text{total}}{V_T} = \frac{14meq}{(10 + 5 + 20)mL} = 0.4N$$

108. A solution of $KMnO_4$ is reduced to MnO_2 and has a normality of 2.75N. calculate

a) The equivalent mas of $KMnO_4$

b) The molarity of the solution

Solution

a) $KMnO_2 + 3e^- \longrightarrow MnO_2$

$$Z = 3\frac{eq}{mol}$$

$$E = \frac{M_{wt}}{Z} = \frac{158g/mol}{3eq/mol} = 52.7g/eq$$

b) $$M = \frac{N}{Z} = \frac{2.75eq/L}{3eq/mol} = 0.92M$$

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109. The density of a 2M solution of weak acid HA is 1.02/ml. Calculate the molality of the solution. (MHA = 48g/mol)

Solution

$$m_{solution} = \rho \times V = 1.02g/\cancel{ml} \times 1000\cancel{ml} = 1020g$$

$$m_{HA} = n \times M_{wt} = 2mol \times 48\frac{g}{mol} = 96g$$

$$m_{solvent} = m_{solution} - m_{HA} = 1020 - 96 = 924g$$

$$= 0.924kg$$

$$m = \frac{n}{kg\,solvent} = \frac{2mol}{0.924kg} = 2.6m$$

or you may use the formula below

$$m = \frac{1000M}{1000\rho - M \cdot M_{solute}} \quad M - molarity$$

m - molatrity M_{solute} - molar mass of solute

ρ – density

$$\Rightarrow m = \frac{1000 \times 2}{1000 \times 1.02 - 2 \times 48} = 2.16m$$

110. What mass of water must be added to 425g of formaldehyde to prepare 40% by mass solution of formaldhyde (formalin)?

Solution

$$\%\ solute = \frac{m_{solution}}{m_{solution} + m_{water}} \times 100$$

$$40 = \frac{425}{425 + m\,water}$$

$$\Rightarrow m_{water} = 637.5g$$

or

Let the mass of solution be x

$$40\%x = 425g$$

$$x = 1032.5g$$

$$m_{water} = 1062.5 - 425 = 637.5g$$

111. How much water is required to dilute 10ml of 10N hydrochloric acid to 0.1N.

Solution

$$N_i V_i = N_f V_f$$

$$V_f = \frac{N_i V_i}{N_f} = \frac{10ml \times 10N}{0.1N} = 1000ml$$

$$VH_2O\ added = V_f - V_i = 1000 - 10 = 990ml$$

112. What volume of 0.25M NaOH solution is required to react with 0.135L of 0.5M H_3PO_4?

Solution

$$n(H_3PO_4) = 0.5\frac{mol}{\cancel{L}} \times 0.135\cancel{L} = 0.675mol$$

$$3NaOH + H_3PO_4 \rightarrow Na_3PO_4 + 3H_2O$$

$$\frac{x}{3} = \frac{0.0675mol}{1}$$

$$x = 3 \times 0.0675mol = 0.2025mol = n(NaOH)$$

$$V = \frac{n}{M} = \frac{0.2025mol}{0.25mol/L} = 0.81L$$

113. What molarity of 7.5ml Na_2SO_4 solution is needed to precipitate barium as $BaSO_4$ from 12.5ml of 0.15M $Ba(NO_3)_2$ solution?

Solution

$$Na_2SO_4\,(eq) + Ba(NO_3)\,(aq) \longrightarrow 2Na\,NO_3\,(aq) + BaSO_4(s)$$

$$n(Ba(NO_3)_2) = M \times V = 12.5ml \times 0.15\,mol/L$$

$$= 1.875mmol$$

since the coefficent ratio of Na_2SO_4 to $Ba(NO_3)_2$ is 1 to 1, $n(Na_2SO_4) = 1.875mmol$

$$M = \frac{n}{V} = \frac{1.875\ mmol}{7.5ml} = 0.25M$$

114. When equeous solution of K_3PO_4 reacts with $Ca(NO_3)_2$,

$Ca_3(PO_4)_2$ precipitates.

a) Write molecular equation

b) Write total ionic equation

c) Identify spectator ion /s/

d) Write net ionic equation

Solution

a) $2K_3PO_4(aq) + 3Ca(NO_3)_2(aq) \rightarrow Ca_3(PO_4)_2(s) + 6KNO_3(aq)$

b) $6K^+(aq) + 2PO^{3-}_4(aq) + 3Ca^{2+}(aq) + 6NO^-_3(aq) \rightarrow 6K^+(aq) + 6NO^-_3(aq) + Ca_3(PO_4)_2(s)$

c) K^+ and NO^-_3

d) $2PO^{3-}_4(aq) + 3Ca^{2+}(aq) \rightarrow Ca_3(PO_4)_2(s)$

Supplementary Exercise

I. True or False Items

1. In a mixture, component keeps its property and separation of components is easy.
2. Suspensions are homogeneous mixture in which the particles settle down.
3. The particles in colloids are bigger in size than those of suspensions.
4. In the solution process, separation of solute and sovent particles absorb energy.
5. The solubility of oxygen in water is due to dipole - force interaction.
6. A mixture of helium and argon forms an ideal solution with positive heat of solution.
7. Supersaturated solution contains more than equilibrium concentration.
8. For exothermic mixing, increasing temperature increases solubility of solute.
9. An association colloid is consists of micelles.
10. In electrolytic solutions, the charge carriers are ions.

II. Multiple Choice items.

11. All of the following are liquid solutions. **EXCEPT:**
A) Soda water C) Brass
B) Soft drink D) Sugar solution

12. Which of the following are immiscible
A) C_2H_5OH and CH_3OH C) CCl_4 and C_6H_6
B) H_2O and CH_3OH D) C_7H_8 and H_2O

13. The solvent in an aqueous solution is
A) Acetone B) Water C) Ethanol D) Chloroform

14. What type of solute - solvent attraction is responsible for the solubility of bromine in acetone?
A) London force C) Dipole - force
B) Hydrogen bond D) Ion - dipole force

15. During dilution, which of the following remains same?
A) Volume of solution C) Amount of solvent
B) Concentration of solution D) Amount of solute

16. When the molarity of sulphuric acid is doubled, its normality is:
A) Double B) Halved C) Quadsupled D) Unchanged

17. Concrete contains all of the following. **EXCEPT**
A) Sand B) Cement C) Paint D) Water

18. " The solubility of a gas is directly proportional to its partial pressure above the solution". This statement is
A) Dalton's law C) Raoult's law
B) Henry's law D) Thomson's law

19. Which gas has the highest Henry law constant at 25°C?
A) He B) H_2 C) N_2 D) O_2

20. Sulpheric acid reacts with potassium hydroxide to form potassium

hydrogen sulphate and water. The equivalent mass of the acids is

A) 98g/eq B) 49 g/eq C) 24.5 g/eq D) 32.6 g/eq

21. A bleaching solution contains 5.2 mass percent NaClO. What is the mass of NaClO in a bottle containing 2500g of bleaching solution?

A) 130g B) 90g C) 480g D) 520g

22. What is the normality of a 1M solution of H_3PO_4 given the following rection?

$H_3PO_4 + 2NaOH \rightarrow Na_2HPO_4 + 2H_2O$

A) 3N B) 2N C) 0.5N D) 1N

23. If 20ml of 0.5N salt solution is diluted to 1L, What is the new concentration?

A) 0.1N B) 0.001N C) 0.01N D) 10N

24. Which of the following has the highest concentration of nitrate ion (NO_3^-)?

A) 0.5M $NaNO_3$ C) 0.3M $Al(NO_3)_3$

B) 0.2M $Ca(NO_3)_2$ D) 0.4M $Ba(NO_3)_2$

25. The number of moles of hydrochloric acid present in 3L of 0.5M aqueous solution is

A) 1.5mol B) 15mol C) 0.25mol D) 3.5mol

26. The molality of 2% (W/W) aqueous solution of sodiumchloride is

A) 0.1m B) 1m C) 0.7m D) 0.35m

27. When saturated solution of sodium chloride is heated, it becomes

A) Unsaturated C) Remains saturated

B) Supersaturated D) Diluted

28. What volume of 0.8M solution contains 0.1 mole of the solute.

A) 100ml B) 500ml C) 125 ml D) 62.5 ml

29. The molarity of 20% solution of sulphuric acid is 2.55M. The density of the solution is:-

A) 1.25g/cm³ C) 0.125 g/L

B) 2.55 g/cm³ D) Upredic table

30. The weight of pure NaOH required to prepare 250cm³ of 0.1N solution is:

A) 4g B) 1g C) 2g D) 10g

III. Short answer Items

31. A solution that contains as much less solute as will dissolve at a given temperature is said to be _____

32. A solution that is 10% by mass sucrose contains 10g of sucrose in every ___ grams solution

33. _____ is a concentrated solution that will be diluted to some lower concentration practical use.

34. _____ is scattering of a beam of light by colloidal particles.

35. A colloid can be _____ , _____ , _____ , _____ , or _____ .

IV Workout Items

36. Calculate the concentration of ions in

a) 0.4M $Co(NO_3)_2$

b) 1M $Al_2(SO_4)_3$

37. What mass of potassium permanganate is needed to prepare a 0.125m $KMnO_4$ solution with 500g of water?

38. What volume of 0.50N H_2SO_4 is required to react exactly with 0.25L of 0.80N KOH?

39. 50g of some dental cream has 0.02g of fluoride concentration. The concentration of fluorine in ppm is _____

40. The mole fraction of solute in 2.5m aqueous solution is _____

Answer

Part I

1. True	3. False	5. False	7. True	9. True
2. False	4. True	6. False	8. False	10. True

Part II

11. C	15. D	19. A	23. C	27. B
12. D	16. C	20. A	24. C	28. C
13. B	17. C	21. A	25. A	29. A
14. A	18. B	22. B	26. D	30. B

Part III

31. Unsaturated solution

32. 100

33. Stock solution

34. Tyndall effect

35. gel, emulsion, sol, foam and aerosol

Part IV

36.

a) $[Co^{2+}] = 0.4M$

$[NO_3^-] = 0.8M$

b) $[Al^{3+}] = 2M$

$[SO_4^{2-}] = 3M$

37. 9.875g

38. 0.4L

39. 400ppm

40. 0.043

UNIT 8

Important Inorganic Compounds

Introduction

- ☑ **Inorganic compounds** are also called **minerals compounds.**
- ☑ They are compounds with out the element carbon. However, oxides of carbon, carbonates, hydrogen carbonates, cyanides, carbides, cyanates are inorganic compounds.
- ☑ They occur mainly as oxides, sulphates, carbonates, sulphides, phosphates, nitrates, silicates and chlorides.
- ☑ They melt or vapourize unlike organic compounds.
- ☑ They are generally found in non-living things.
- ☑ Minerals and rocks are made up of inorganic compounds.
- ☑ Inorganic compounds are classified into four.

1. Oxides 2. Acids 3. Bases 4. Salts

8.1 Oxides

- ☑ Oxides are binary compounds containing oxygen
- ☑ Oxygen reacts almost with all elements to form oxide except noble gases, and noble metals such as gold, platinum, palladium

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☑ **Binary compounds** are compounds formed from two types of elements.

☑ When metals react with oxygen, the resulting oxides are **metallic oxides.**

Metal + oxygen → Metallic oxide

Example

Magnesium + oxygen → magnesium oxide

$2Mg + O_2 \rightarrow 2MgO$

The combination of non-metals with oxygen results **non-metallic oxides.**

Non - metal + oxygen → Non - metallic oxide

Example

Sulphur + oxygen → Sulphur dioxide

$S + O_2 \rightarrow SO_2$

Illustrative Examples

1. Which one of the following is not considered as inorganic compound?

A) Urea B) Water C) Salt D) baking soda

Solution

Urea is the first synthetic organic compound.

Answer: A

2. Inorganic compounds are classified into _____

A) Two B) Three C) Four D) Five

Answer: C

3. All of the following are oxides. **EXCEPT**

A) H_2O B) HNO_3 C) CO_2 D) Al_2O_3

Solution

Oxides are made up of two types of elements. One of the element must be oxygen.

Answer: B

4. Which one of the following is not non - metallic oxide?

A) P_2O_5 B) N_2O_5 C) SO_3 D) CaO

Solution

Calcium is metallic element. Thus, CaO is metallic oxide.

Answer: D

5. Classify the following compounds as organic or inorganic compounds.

a) Enzymes d) Oil g) Proteins

b) Limestone e) Fats h) Carbohydrates

c) Sulphuric acid f) Washing soda i) Sudiumchloride

Solution

➥ a, d, e, g and h are organic compounds.

➥ b, c, f and i are inorganic compounds.

6. List some metals and non - metals that do not react-with oxygen.

Solution

Metals	Non - metals
➥ Gold (Au)	➥ Helium (He)
➥ Platinum (Pt)	➥ Neon (Ne)
➥ Palladium (Pd)	➥ Argon (Ar)
➥ Osmium (Os)	➥ Krypton (Kr)
➥ Rutherium (Ru)	

7. Classify the following as metallic or non - metallic oxide.

a) SO_3 c) V_2O_5 e) N_2O_5

b) K_2O d) CO_2 f) NiO_2

Solution

➥ a, d and e are non-metallic oxides

➥ b,c and f are metallic oxides

Classification of Oxides

Based on their properties, oxides are classified into five.

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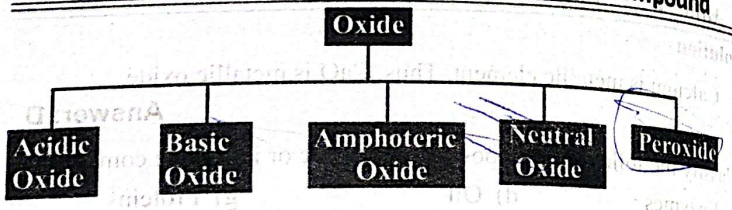

Fig 8.1 The Five Types of Oxides

1. Acidic oxides /Acid anhydrides/

- They are oxides that have acidic character
- Most of them are non-metallic oxides

Example

- CO_2 - Carbondioxide
- SO_2 - Sulphur dioxide
- SO_3 - Sulphur trioxide
- NO_2 - Nitrogen dioxide
- N_2O_3 - Dinitrogen trioxide
- N_2O_5 - Nitrogen pentoxide
- P_2O_3 - Phospharus trioxide
- P_2O_5 - Phosphons pentoxide

Oxides of metalloids are acids oxides

Example

- B_2O_3 - Boron oxide
- $Si O_2$ - Silicondioxide
- GeO_2 - Germenium oxide

Few metallic oxides are acidic oxides.

- CrO_3 - Chromium (VI) oxide
- Mn_2O_7 - Manganes (VII) oxide

Chemical Properties of Acidic Oxides

1. **Acidic oxides are also known as acid anhydrdes, because they form acids when dissolved in water**

Acidic oxide + water → Acid

Example

$SO_3 + H_2O \rightarrow H_2SO_4$

Note : To determine the acid anhydride of an acid, use the simple formula below:

Acid anhydride = Acid - water

- First, balance the number of hydrogen atoms in water and acid.

Example

Determine the acid anhydride of:

a) H_2SO_3 b) HNO_3

Solution

a) Acid an hydride = Acid - H_2O

= H_2SO_3 - H_2O (H is already balanced)

⇒ 2H – 2H = Zero

S - Zero = S

3O – O = O_2

⇒ Thus, the acid anhydride of H_2SO_3 is SO_2

b) Acid anhydride = Acid - H_2O

= HNO_3 - H_2O (H - is not balanced)

= $2HNO_3$ - H_2O

⇒ 2H - 2H = Zero

⇒ 2N - Zero = N_2

⇒ 6O - O = O_5

Thus, the acid anhydride of HNO_3 is N_2O_5

2. **Acidic Oxides React with Basic Oxides**

Acidic oxide + Basic oxide → salt

Example

$SO_2 + CaO \rightarrow CaSO_3$ (calcium sulphite)

$CO_2 + MgO \rightarrow MgCO_3$ (Magnesium carbonate)

3. Acidic Oxides React with Bases.

Acidic oxide + Base → Salt + water

Example

$SO_2 + Ca(OH)_2 \rightarrow CaSO_4 + H_2O$

$CO_2 + 2KOH \rightarrow K_2CO_3 + H_2O$

Note: The reaction between acidic oxide and base, acid and base or acid and basic oxide to yield salt and water is known as **neutralitation reaction.**

2. Basic Oxides (Basic anhydrides)

- They have basic character.
- All basic oxides are metallic oxides but not the reverse

Example of basis oxides

- Li_2O - Lithium oxide
- Na_2O - sodium oxide
- K_2O - Potassium oxide
- MgO - Magnesium oxde
- CaO - Calcium oxide
- SrO - Strontium oxide
- BaO - Barium oxide
- FeO - Iron (II) oxide.

Chamical Properties of Basic Oxides

1. They React with Water to yield Bases.

Basic oxide + water → Base

Example

$K_2O + H_2O \rightarrow 2KOH$ (potassium hydroxide)

$CaO + H_2O \rightarrow Ca(OH)_2$ (calciumhydroxide)

Note

Basic anhydride = Base - water

Example

Identify the basic anhydride of $Ca(OH)_2$

Solution

Basic anhydride = $Ca(OH)_2 - H_2O = CaO$

Thus, CaO is basic anhydride of $Ca(OH)_2$

2. Basic oxides react with acidic oxides to give salts.

Basic oxide + Acidic oxide → salt

Example

$Na_2O + CO_2 \rightarrow Na_2CO_3$ (soldium carbonate)

3. Basic oxides react with acids to give salt and water

Basic oxide + acid → salt + water

Example

$BaO + H_2SO_4 \rightarrow BaSO_4 + H_2O$

Bariumsulphate

$CuO + 2HCl \rightarrow CuCl_2 + H_2O$

Copper (II) Choride

Identification of Acidic and Basic Oxides

- Acidic and basic oxides are identified by using **indicators.**
- **Indicators** are substances that show color change in acidic and basic solutions.

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Table 8.1 The Colors of Some indicators in Acidic and Basic Solutions

Indicator	Color in acidic solution	Color basic solution
Litmuspaper	Red	Blue
Phenolphthalien	Colorless	Pink (Red)
Methylorange	Red	Yellow
Universal indicator	Yellow, orange, red	Bule, purple

Illustrative Example

8. Which one is an acidic oxide?

A) CuO B) Fe_2O_3 C) SiO_2 D) Na_2O

Answer: C

9. What compound does "X" represent in the chemical equation below?

$x + H_2O \rightarrow HNO_3 + HNO_2$

A) NO_2 B) NO C) N_2O_5 D) N_2O_3

Solution

Nitrogen dioxide (NO_2) is commonly known as mixed acid anhydride because it forms two types of acids (HNO_3) and (HNO_2) when treated with water.

Answer: A

10. Which of the following is not basic oxide?

A) K_2O B) CaO C) Al_2O_3 D) BaO

Answer: C

11. Identify the acid anhydride of

a) H_2SO_4 c) HNO_3 e) H_3PO_3

b) H_2CO_3 d) H_3PO_4 f) HNO_2

Solution

a) $\Rightarrow H_2SO_4 - H_2O = SO_3$

b) $\Rightarrow H_2CO_3 - H_2O = CO_2$

c) $\Rightarrow 2HNO_3 - H_2O = N_2O_5$

d) $\Rightarrow 2H_3PO_4 - 3H_2O = P_2O_5$

e) $\Rightarrow 2H_3PO_3 - 3H_2O = P_2O_3$

f) $\Rightarrow 2HNO_2 - H_2O = N_2O_3$

12. Identify the basic anhydride of the following bases.

a) LiOH c) $Mg(OH)_2$ e) $Sr(OH)_2$

b) KOH d) $Ca(OH)_2$ f) $Fe(OH)_3$

Solution

a) Basic anhydride = base - H_2O

$= 2LiOH - H_2O = Li_2O$

b) Basic anhydride = $2KOH - H_2O = K_2O$

c) Basic anhydride = $Mg(OH)_2 - H_2O = MgO$

d) Basic anhydride = $Ca(OH)_2 - H_2O = CaO$

e) Basic anhydride = $Sr(OH)_2 - H_2O = SrO$

f) Basic anhydride = $2Fe(OH)_3 - 3H_2O = Fe_2O_3$

13. Complete and balance the following

a) $BaO + P_4O_{10} \rightarrow$ c) $CaO + H_2S \rightarrow$

b) $SiO_2 + MgO \rightarrow$ d) $SO_3 + LiOH \rightarrow$

Solution

a) $6BaO + P_4O_{10} \rightarrow 2Ba_3(PO_4)_2$

b) $SiO_2 + MgO \rightarrow MgSiO_3$

c) $CaO + H_2S \rightarrow CaS + H_2O$

d) $SO_3 + 2LiOH \rightarrow Li_2SO_4 + H_2O$

14. When magneium is burnt with air, white ash is formed. The resulting product turns red litmus to blue. What type of product is it?

Answer: Basic oxide

3. Amphoteric Oxides

- They display both acidic and basic properties
- They react with acids and bases.

Example

Al_2O_3 - Aluminum oxide

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ZnO - Zinc oxide

PbO - Lead (II) oxide

PbO - Lead (IV) oxide

SnO - Tin (II) oxide

SnO_2 - Tin (IV) ixide

Chemical properties of Amphoteric Oxides

1. They React with Acids

Amphoteric oxide + acid → Salt + water

Example

$Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$

2. They React with Bases

Amphoteric oxide + base → salt + water

Example

$Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$

Note: Acidic oxides, basic oxides and amphoteric oxides are salt forming oxides.

4. Neutral Oxides

They do not have basic and acidic characters.

Example

H_2O - water

NO - Nitric oxide

CO - carbon monoxide

N_2O - Nitrous oxide (laughing gas)

5. Peroxides

These are oxides in which the two oxygen atoms are linked to each other by —O—O—link or bond.

The oxidation number of oxygen in perioxides is **-1** but it is **-2** in the above oxides.

Example

H_2O_2 - Hydrogen peroxide

Na_2O_2 - Sodium peroxide

K_2O_2 - Potassium peroxide

CaO_2 - Calcium peroxide

SrO_2 - Strontium peroxde

BaO_2 - Barium peroxide etc.

Chemical Properties of Perxides

1. **Peroxides react with substances by losing oxygen.**
   - They are strong oxidizing agent

Example

$HgS + 4H_2O_2 \rightarrow HgSO_4 + 4H_2O$

2. **Peroxides react with dilute acids to form salt and hydrogen peroxide.**

Example

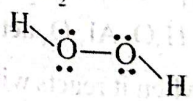
$K_2O_2 + 2HCl \rightarrow 2KCl + H_2O_2$

Hydrogen Peroxide (H_2O_2)

- It is strong oxidizing and bleaching agent.
- It is unstable, decomposing readily to water and oxygen.

$2H_2O_2 \rightarrow 2H_2O + O_2$

- Its structure:



Illustrative Example

15. An oxide in which the two oxygen atoms are linked to each other y —O—O— link and with atoms of other elements is known as:

A) An acidic oxide

B) A basic oxide

C) An amphoteric oxide

D) A peroxide

Answer: D

16. Which of the following pair of oxides are amphoteric?

A) PbO and CaO_2 C) Al_2O_3 and Na_2O_2

B) Sn O and H_2O_2 D) ZnO and PbO

Answer: D

17. All of the following are neutral oxides. EXCEPT

A) NO_2 B) N_2O C) NO D) CO

Solution

NO_2 is mixed acid anhydride.

Answer: A

18. Which of the following is true about oxides?

A) Oxides are ternary compounds containing oxygen

B) All basic oxides are metallic oxides.

C) Neutral oxides react with acids and bases.

D) All non - metallic oxides are acidic oxides.

Answer: B

19. Write chemical equations to show the amphoteric oxides of Al_2O_3 with

a) HNO_3 b) KOH

Solution

a) $Al_2O_3 + 6HNO_3 \rightarrow 2Al(NO_3)_3 + 3H_2O$, Al_2O_3 acts as base

b) $Al_2O_3 + 2KOH \rightarrow 2KAlO_2 + H_2O$, Al_2O_3 acts as an acid

20. The type of oxide that forms salt when it reacts with acid and base is

Answer: Amphoteric oxide

8.2 Acids and Bases

Acids and bases are important substances. Their definitions were given by different scientists as follows:

1. Arrhenius Definition

➥ **Acid** is a substance that yields hydrogen ion (H^+) when dissolved in water

➥ It increases concentration of hydrogen ion when added to water.

Example

HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_4$ etc.

$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$

➥ **Base** is a substance that yields hydroxide (OH^-) ion when dissolved in water.

➥ It increases the concentration of hydroxide ion when disolved in water.

Example of Arrhenius bases:

NH_3, NaOH, KOH, $Ba(OH)_2$, $Ca(OH)_2$ etc.

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

The acid - base reaction for Arrhenius defintion is:

Acid + base → Salt + water

Limitation of Arrheniusn Difinition:

The definition is restricted to aqueous phase He didn't explain the acid - base definition in gaseous and other phases.

Example

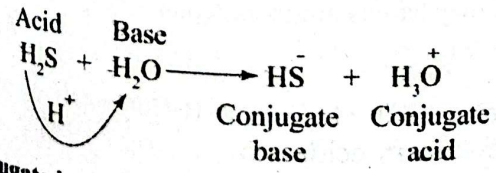
$HCl(g) + NH_3(g) \rightarrow NH_4Cl(s)$

2. Bronsted - Lowry definition

➥ Acid is proton (H^+) donor

➥ Base is proton acceptor

Example



➥ **Conjugate base** is formed when an acid donates a proton.

➥ **Conjugate acid** is formed when a base accepts a proton

Conjugate acid (CA) = Base + H^+

Conjugate base (CB) = acid - H^+

Example

Determine the CA and CB of H_2O

Solution

CA = base + H^+ = $H_2O + H^+ = H_3O^+$

CB = acid - H^+ = $H_2O - H^+ = OH^-$

3. Lewis Definition

➥ **Acid** is electron - pair acceptor

➥ **Base** is electron - pair donor

Illustrative Examples

21. Arrhenius acid is a substance that:

A) releases hydronium ion in water

B) releases hydroxide ion in water

C) accepts a proton from a base

D) donates electron pair to a base

Answer: A

22. According to Bronsted - Lowry definition, a base is:

A) electron donor acceptor C) Proton donor

B) electron pair donor D) Proton acceptor

Solution

Bronsted - Lowy base is proton acceptor

Answer: D

23. Give the reaction: $HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + OH^-$.
The Bronsted - Lowry acids are :

A) H_2O and H_2CO_3 B) HCO_3^- and CO_3^{2-}

C) H_2O and OH^- D) HCO_3^- and H_2CO_3

Solution

Bronsted - Lowry acid is proton donor. For the forward reaction, H_2O and for the backward reaction, H_2CO_3 are acids.

Answer: A

24. Among the following, identify a substance that is capable of donating a pair of electrons.

A) An arrhenius acid C) a Lewis base

B) a Lewis acid D) An arrhenius base

Answer: C

25. Determine the CA of

a) P c) HS^- e) CN^-

b) CO_3^{2-} d) HPO_4^{2-} f) NH_3

Solution

a) CA = P + H^+ = PH^+

b) CA = CO^{2-}_3 + H^+ = HCO_3^-

c) CA = HS^- + H^+ = H_2S

d) CA = $H_2PO_4^-$ + H^+ = $H_2PO_4^-$

e) CA = CN^- + H^+ = HCN

f) CA = NH_3 + H^+ = NH^+_4

26. Identify the CB of:

a) H_3PO_4 c) HCO_3^- e) NH_3

b) $HClO_3$ d) $HClO_3$ f) HS^-

Solution

a) CB = H_3PO_4 - H^+ = $H_2PO_4^-$

b) CB = $HClO_3$ - H^+ = ClO^-_3

c) CB = HCO^-_3 - H^+ = CO_3^{2-}

d) CB = $N_2H_5^+$ - H^+ = N_2H_4

e) CB = NH_3 - H^+ = NH_2^-

f) CB = HS^- - H^+ = S^{2-}

27. Write the limitation of Bronsted - Lowry of definition

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Solution

Their definition is restricted to proton transfer

28. The more general and modern definition of acid and base is that of ____

Answer: Lewis definition

General Properties of Acids and Bases

Acids and bases have the following general properties.

Table 8.2 General properties of Acids and Bases

Acids	Bases
1. They have sour taste	1. They have bitter taste
2. They release hydrogen ion in water	2. They relese hydroxide ion in water
3. They change color of indicator	3. They change color of indicators
Example	**Example**
➥ Blute litmus to red	☑ red limtmus to blue
➥ Phenolphthalien to colorles	☑ Phenolphthalien to red or pink
➥ Methyl orange to red	☑ Methyl orange to yellow
➥ Universal indicator to yellow, organe, red	☑ Universal indicntor to blue, purple
4. Dilute acids react with active metals to form salt and hydrogen gas.	4. Bases react with acids
Example	Base + acid → salt + water
$Zn + 2HCl \rightarrow ZnCl_2 + H_2$	**Example**
5. Acids react with bases and basic oxides.	$NaOH + HCl \rightarrow NaCl + H_2O$
Acids + base/basic oxide → salt + water.	5. Bases react with acidic oxides
Example	Base + acidic oxide → salt + water.
$HCl + KOH \rightarrow KCl \rightarrow H_2O$	
$2HCl + K_2O \rightarrow 2KCl + H_2O$	
6. Acids react with carbonates and hydrogen carbonates.	

Acid + carbonate/hydrogen carbonare → Salt + carbondioxide + water	**Example**
Example	$Ca(OH)_2 + SO_3 \rightarrow CaSO_4 + H_2O$
$CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + O_2 + H_2O$	8. Bases are electroyes.
6. Acids are electrolytes	9. pH > 7
7. pH < 7	

Illustrative Examples

29. Which of the following is not true about acids?

A) They change the color of litmus to red

B) All acids release hydrogen ion in water

C) They are electrolytes

D) They react with active metals

Answer: B

30. Which of the following gas is formed when dilute acids react with active metals?

A) H_2 B) CO_2 C) O_2 D) N_2

Solution

Active metal + acid → salt + hydrogen gas

Answer: A

31. Which of the following is the common property of acid and base?

A) Both react with carbonates

B) Both have sour taste

C) Both release hydronium ion in water

D) Both are electrolytes.

Answer: D

32. All of the following are the properties of bases **EXCEPT**

A) They are slippery to touch

B) They react with both acid and acidic oxides

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C) Aqueous solutions of bases are electrolytes.

D) They change phenolphthalien to blue.

Solution

The color of phenolphthalien in basic solution is pink.

Answer: D

33. List some metals that do not react with acids

Solution

Gold, silver, platinum, palladium, osmium.

34. Complete and balance the following

a) $Al + HCl \rightarrow$

b) $H_2SO_4 + NaOH \rightarrow$

c) $HNO_3 + CaO \rightarrow$

d) $HBr + MgCO_3 \rightarrow$

e) $HClO_4 + KHCO_3 \rightarrow$

Solution

a) $2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$

b) $H_2SO_4 + K_2O \rightarrow K_2SO_4 + H_2O$

c) $2HNO_3 + CaO \rightarrow Ca(NO_3)_2 + H_2O$

d) $2HBr + MgCO_3 \rightarrow MgBr_2 + CO_2 + H_2O$

e) $HClO_4 + KHCO_3 \rightarrow KClO_4 + CO_2 + H_2O$

35. Why acids and bases transmit electricity in aqueous state?

Solution

Because they produce ions that carry charge

Classification of Acids

There are different ways used to classify acids

1. Based on the Types of Elements They Contain:

Binary acids are acids formed from two types of elements.

Example

HF - Hydrofluoric acid

HCl - Hydrochloric acid

HBr - Hydrobromic acid

HI - Hydroiodic acid

H_2S - Hydrosulphuric acid etc.

2. Ternary acids are formed from three types of elements.

Example

HNO_3 - Nitric acid

H_2SO_4 - Sulphuric acid

$HClO_4$ - Perchloric acid

HCN - Hydrocyanic acid

H_3PO_4 - Phospheric acid etc.

I. Based on the Number of Ionizable Hydrogen

1. **Monoprotic acids** - contain one ionizable hydrogen per formula.

Example

HCl, HI, HF, HBr, HNO_3, HNO_2, $HClO_4$ etc.

$HCl \rightarrow H^+ + Cl^-$

2. **Diprotic acids** - have two ionizable hydrogens per formula

Example

H_2CO_3, H_2SO_4, $H_2C_2O_4$, H_2SO_3 etc

$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$

3. **Triprotic acids** - possess three ionizable hydrogens per formula.

Example

H_3PO_4, H_3AsO_4 etc

$H_3PO_4 \rightarrow 3H^+ + PO^{3-}_4$

Note: Acids that have two or more ionizable hydrogens per formula are called **polyprotic acids.**

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II. Based on Degree of Ionization or Dissociation

1. Strong acids are strong elecrolytes which, for most practical purposes, are assumed to ionize completely in water.

- They exist almost 100% in water in the form of ions.
- Their dissociation is represented by single arrow (→).

Example

- HBr, HCl, HI, HNO_3, $HClO_4$, H_2SO_4

$$HBr \longrightarrow H^+ + Br^-$$

2. Weak acids ionize only to a limited extent in water.

- Their ionization reaction is represented by double arrows ($\rightleftharpoons$)

Example

HF, H_2S, HCN, H_2CO_3, H_2SO_3, HNO_2, CH_3COOH etc

$$HF \rightleftharpoons H^+ + F^-$$

III. Based on the Amount of Acids in Water

1. Dilute acid contains less solute (acid) and more solvent

A dilute acid can be strong or weak acid

2. Concentrated acid is a solution of acid with more acid and less solvent.

- It can be strong or weak.

Example

- A 1% H_2CO_3 is dilute weak acid because out 100, only 1% acid and 99% is water.
- A 98% H_2SO_4 is concentrated strong acid since out 100%, 98% is acid and 2% is water.

Note: The most important concentration unit is molarity.

$$\text{Molarity} = \frac{\text{Number of mole of solute}}{\text{Volume of solution in liter}}$$

$$\boxed{M = \frac{n}{V}} \qquad \boxed{n = M \times V}$$

The unit of molarity moleper liter (mol/L) or molar (M): $\boxed{1M = 1mol/L}$

Example

1. Calculate the molarity of a solution obtained by mixing 0.01mol H_2SO_4 and 500ml of solution.

Solution

1000ml = 1L n = 0.01mol

500ml = ? M = ?

$$V = \frac{500\,ml \times 1L}{1000\,ml} = 0.5L$$

$$M = \frac{n}{V} = \frac{0.01mol}{0.5L} = 0.02mol/L$$

$$= 0.02M$$

2. Calculate the number of moles of 2M of 3L hydrochloric acid solution.

Solution

$$n = ?\ M = 2M = 2\frac{mol}{L},\ V = 3L$$

$$n = M \times V = 2\frac{mol}{L} \times 3L = 6mol$$

IV. Based on Source, Acids are Classified as Inorganic or Organic Acids

1. **Inorganic acids** are also called mineral acids because they are inorganic compounds.

Example

HNO_3, HCl, H_2SO_4, H_3PO_4 etc.

2. **Organic acids** are also known as carboxylic acids, because they all

contain carboxyl group (-COOH)

Example

CH_3COOH, HCOOH, HOOC - COOH etc.

Organic acids are generally weak acids

Illustrative Example

36. Which one of the following is ternary acid?

A) $HClO_4$ B) H_2S C) H_2Se D) HI

Answer: A

37. The strength of an acid can be determined by

A) Its degree of ionization

B) Its number oxygen atoms.

C) Its number of hydrogen atoms

D) Its molar concentration

Solution

The strength of an acid is determined by its degree of ionization or dissociation.

Answer: A

38. A certain acid that was purchased from a super market ionizes almost completely in water. Then, it is a:

A) Dilute acid C) Strong acid

B) Concentrated acid D) Weak acid

Answer: C

39. Which of the following statement is not true?

A) A dilute solution of HCl is dissolved in a smaller volume of water

B) A concentrated solution of HCl is dissolved in smaller volume of water

C) A concentrated acid has relatively large amount of solute dissolved in the solvent

D) A diluted acid solution has a relatively small amount of solute dissolved in the solvent.

Answer: A

40. Classify the following acids as strong or weak acid.

a) HF c) H_2SO_4 e) HNO_2 g) $HClO_4$

b) HCOOH d) H_4SO_3 f) HNO_3 h) $HClO_2$

Solution

- a, b,d, e and h are weak acids.
- c,f and g are strong acids

41. Categorize the following acids as monoprotic, diprotic or triprotic acid.

a) H_2S c) H_3PO_4 e) $HBrO_3$

b) HClO d) H_2CO_3 f) H_3AsO_4

Solution

- a and d are diprotic acids
- b and e are monoprotic acids
- c and f are triprotic acids

42. Write three example of organic acids

Solution

- CH_3COOH - Acetic acid
- HCOOH - formic acid
- HOOC - COOH - Oxalic acid

43. Explain the following

a) 99% nitric acid b) 1% nitric acid

Solution

a) 99% HNO_3 and 1% water

b) 1% HNO_3 and 99% water

44. Calculate the molarity of 24.5g of sulphuric acid in 400ml of sulphric acid solution.

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Solution

$V = 400ml = 0.4L$

$$n = \frac{m}{M_{wt}} = \frac{24.5g}{98g/mol} = 0.25mol$$

$$M = \frac{n}{V} = \frac{0.25mol}{0.4L} = 0.625M$$

45. What is the mass of a 2M H_2SO_4 in a 250ml of solution?

Solution

$M = 2M, \quad V = 250ml = 0.25L$

$M_{wt} = 98g/mol, \quad m = ?$

$$M = \frac{m}{M_{wt} \times V}$$

$$m = M \times M_{wt} \times V = 2\frac{mol}{\cancel{L}} \times 98\frac{g}{mol} \times 0.25\cancel{L} = 49g$$

46. How many milliliters of solution are required to prepare 0.02mol of 2.5M NaCl solution?

Solution

$$n = 0.2mol, \; M = 2.5\frac{mol}{L}, \; V = ?$$

$$V = \frac{n}{M} = \frac{0.02mol}{2.5\frac{mol}{L}} = 0.008L = 8mL$$

Classification of Bases

I. Based on Degree of Ionization

1. Strong bases ionize almost completely in water

Example

LiOH - Lithium hydroxide

NaOH - Sodium hydroxide

KOH - Potassium hydroxide

$Ca(OH)_2$ - Calcium hydroxide

$Sr(OH)_2$ - Strontium hydroxide

$Ba(OH)_2$ - Barium hydroxide etc.

$LiOH \rightarrow Li^+ + OH^-$

2. Weak bases ionize slightly when dissolved in water.

Example

NH_3 - Ammonia

$Mg(OH)_2$ - Magnesium hydroxide

$Al(OH)_3$ - Aluminum hydroxide

$Fe(OH)_2$ - Iron (II) oxide

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

II. Based on the Amount of Base in Water:

1. Concentrated base - contains relatively large amount or percentage of base.

2. Dilute base - has low a mount of base compared to the quantity of solvent

Example

- 98% NaOH solution is concentrated solution of NaOH
- 2% NaOH solution is diluted solution of NaOH.

III. Based on the Number of Hydroxyl Group they Contain

1. Monobasic base - has one hydroxyl group

Example

KOH, LiOH, NaOH, RbOH, NH_4OH etc

2. Dibasic base - contains two hydroxyl groups

Example

$Mg(OH)_2$, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$, $Fe(OH)_2$ etc.

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3. Tribasic base - possesses three hydroxyl groups

Example

$Al(OH)_3$, $Fe(OH)_3$, etc.

IV. Based on source

1. Inorganic bases are bases that do not have the element carbon.

Example

NaOH, KOH, $Mg(OH)_2$, $Al(OH)_3$ etc.

2. Organic bases are bases containing carbon.

Example

CH_3NH_2 - Methylamine

$CH_3CH_2NH_2$ - Ethylamine

C_5H_5N - Pyridine etc.

Illustrative Example

47. All of the following are strong bases. **EXCEPT**

A) $Ca(OH)_2$ B) $Mg(OH)_2$ C) $Sr(OH)_2$ D) $Ba(OH)_2$

Answer: B

48. Which of the following is not true about the properties of bases?

A) They have bitter taste

B) They are slippery touch

C) They change blue litmus to red

D) They react with acids and acidic oxides

Answer: C

49. Classify the following bases as strong or weak base.

a) CsOH c) $Cu(OH)_2$ e) $Ba(OH)_2$

b) $Mg(OH)_2$ d) $Fe(OH)_2$ f) RbOH

Solution

➥ a, e, and f are strong bases.

➥ b, c and d are weak bases.

50. Generally, organic bases are weak bases. Why?

Solution

Because they ionize very slightly.

Ionization of water

Pure water contains small amount of hydronium and hydroxide ions due to self or auto ionization of water.

$$H_2O(1) + H_2O(1) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

$$KC = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

Multiply both sides by $[H_2O]^2$

$$\Rightarrow Kc[H_2O]^2 = [H_3O^+][OH^-]$$

$$\Rightarrow \boxed{K_w = [H_3O^+][OH^-]},\ K_w = K_c[H_2O]^2$$

K_w - ion product of water

- It is affected by **temperature**

$K_w = 1.0 \times 10^{-14}\ M^2$ at 25^0C

For pure water, $[H_3O^+] = [OH^-]$

$$\Rightarrow K_w = [H_3O^+][OH^-] = 1 \times 10^{-14}\ M^2$$

since $[H_3O^+] = [OH^-]$

$$\Rightarrow [H_3O^+]^2 = 1 \times 10^{-14}\ M^2 \Rightarrow [H_3O^+] = 1 \times 10^{-7}\ M$$

$$\therefore [H_3O^+] = [OH^-] = 1 \times 10^{-7} M$$

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$[H_3O^+] > 1\times10^{-7}M$, $[OH^-] < 1\times10^{-7}M$, $[H_3O^+] > [OH^-]$	acidic solution
$[H_3O^+] < 1\times10^{-7}M$, $[OH^-] > 1\times10^{-7}M$, $[OH^-] > [H_3O^+]$	Basic or alkaline solution
$[H_3O^+] = 1\times10^{-7}M$, $[H_3O^+] = [OH^-]$, $[OH^-] = 1\times10^{-7}M$	Neutral solution

Example

Calculate the $[OH^-]$ or $[OH^-]$ or both at 25^0C

a) $[H_3O^+] = 1\times10^{-6}M$

b) $[HO^-] = 2\times10^{-9}M$

c) $[H_3O^+] = 100[OH^-]$

Solution

a) $[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1\times10^{-14}M^2}{1\times10^{-6}M} = 1\times10^{-8}M$

b) $[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1\times10^{-14}M^2}{2\times10^{-9}M} = 5\times10^{-6}M$

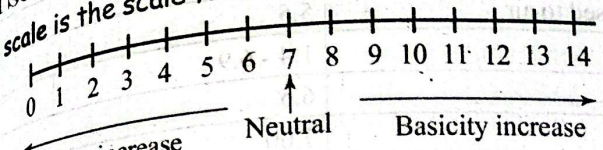
c) $[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1\times10^{-14}M^2}{100[OH^-]}$

$\Rightarrow [OH^-] = \sqrt{\frac{1\times10^{-14}M^2}{100}} = 1\times10^{-8}M$

$\therefore [H_3O^+] = 100[OH^-] = 100\times10^{-8}M = 1\times10^{-6}M$

pH Scale, pH and pOH

PH scale is the scale from 0 to 14.



PH scale is the negative logarith of the concentration of hydrogen or hydronium ion

$[H^+] = [H_3O^+]$

$pH = 10g\frac{1}{[H^+]} = -10g[H^+]$

$[H^+] = [H_3O^+] = 10^{-PH}$

pOH is the negative logarith of the concentration of hydroxide ion.

$pOH = 10g\frac{1}{[OH^-]} = -10g[OH^+]$

$[OH^-] = 10^{-pOH}$

$pH < 7$, $pOH > 7$	acidic
$pH = 7$, $pOH = 7$	Neutral solution
$pH > 7$, $pOH < 7$	Basic solution

Table 8.3 The pH of Some Common Fluids.

Sample	pH
Gastric juice in stomach	1 . 0 - 2.0
Lemon	2.4
Vinegar	3.0

Organge juice	3.5
Urine	1.8 - 7.5
Water exposed to air	5.5
Saliva	1.4 - 6.9
Milk	6.5
Pure water	7.00
Blood	1.35 - 7.45
Tears	7.4
Milk of magnesia	10.6
House hold ammonia	11.5

pH meter is commonly used in the laboratory to determine the pH of solution. Although many pH meters have scales marked from 1 to 14, pH Values can in fact, be less than 1 and greater than 14.

Example

Given 0.1M HCl solution. Calculate

a. $[H^+]$ b. pH c. $[OH^-]$ d. pOH

Solution

a) For strong monoprotic acid,

$\Rightarrow [HCl] = [H^+] = 0.1M$

or used stoichiometry

$HCl \longrightarrow H^+ + Cl^-$

1M 1M

0.1M ?

$\Rightarrow [H^+] = \frac{0.1M \times 1M}{1M} = 0.1M = 10^{-1}M$

b) $pH = -10g[H^+] = 10g\ 10^{-10}$

c) $[OH^+] = \frac{K_w}{[H^+]} = \frac{1 \times 10^{-14}M^2}{10^{-1}M} = 1 \times 10^{-13}M$

d) $pOH = -10g[OH^-] = -10g\ 10^{-13} = 13$

Relationship Between pH and pOH

The relationship pH and pOH can be derived from self - ionization of water at 25°c.

$K_w = [H^+][OH^-] = 1 \times 10^{-14}$

Put longarithm on both sides of the equation

$10g[H^+][OH^-] = 10g\ 10^{-14}$

$10g[H^+] + 10g[OH^-] = 10g10^{-14}$, multiply both sides by -1

$-10g[H^+] + -10g[OH^-] = -10g\ 10^{-14}$

$\boxed{pH + pOH = 14}$

$\boxed{pH = 14 - pOH = 14 - (-10g[OH^-]) = 14 + 10g[OH^-]}$

$\boxed{pOH = 14 - pH = 14 - (-10g[H^+]) = 14 + 10g[H^+]}$

Note:

i. As pH increases, $[H^+]$ decreases but $[OH^-]$ increases and vice versa.

ii. As pOH increases, $[OH^-]$ decreases but $[H^+]$ increases and vice versal.

Example

The pH of a give solution is 2. Determine

a) $[H^+]$ b) pOH c) $[OH^-]$

Solution

a) $[H^+] = 10^{-pH} = 10^{-2}M = 0.01M$

b) $pOH = 14 - pH = 14 - 2 = 12$

c) $[OH^-] = 10^{-pOH} = 10^{-12}M$

Illustrative Examples

51. Given the following table that contains information about three solutions

Solution	pH value
1	10
2	7
3	5

Which list has the solutions placed in placed in order of increasing hydrogen ion concentration?

A) 3,1,2 B) 3,2,1 C) 1,2,3 D) 2,1,3

Solution

The higher the pH of a solution, the lower is the hydrogen ion concentration.

Answer: C

52. Which of the following solutions has the highest hydroxide ion concentration?

A) pH = 5.67 C) pH = 10.59

B) pH = 8.72 D) pH = 13.08

Solution

The higher the pH, the higher is the hydroxide ion concentration.

Answer: D

53. When the pH of a solution is changed from 2 to 1, the concentration of hydronium ions

A) Increases by a factor of 10 C) Increases by a factor of 2

B) Decreases by a factor of 10 D) Decreases by a factor of 2

Solution

Use ratio method to solve the problem

$$\frac{[H^+]_1}{[H^+]_2} = \frac{10^{-pH_1}}{10^{-pH_2}} = \frac{10^{-2}}{10^{-1}} = \frac{1}{10}$$

$$\Rightarrow [H^+]_2 = 10[H^+]_1$$

Thus, the pH increases by afactor 10

Answer: A

54. Which of the following affects the value of ion - product of water, K_w ?

A) Concentration of hydrogen ion

B) Concentration of hydrogen ion

C) Temperature

D) Pressure

Solution

K_w is affected by temperature. As temperature in creases, the value of K_w also increases.

Answer: C

55. What is the pH of a solution that has hydrogen ion concentration 100 times greater than a solution with a pH of 4?

Solution

$pH_1 = 4$, $pH_2 = ?$ if $[H^+]_2 = 100[H^+]_1$

$\Rightarrow [H^+]_1 = 10^{-pH1} = 10^{-4}M$

$[H^+]_2 = 100\,[H^+] = 100 \times 10^{-4}\,M = 10^{-2}M$

$\therefore pH_2 = -10g[H^+]_2 = -10g\ 10^{-2} = 2$

56. Calculate $[H^+]$, pH, [OH], and pOH of 0.5M H_2SO_4 solution. Assume complete ionization.

Solution

$$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$$

1M 2M

0.05M ?

$$\Rightarrow [H^+] = \frac{0.05M \times 2M}{1M} = 0.1M = 10^{-1}M$$

$$\Rightarrow pH = -10g[H^+] = -10g10^{-1} = 1$$

$$\Rightarrow pOH = 14 - pH = 14 - 1 = 13$$

$$\Rightarrow [OH^-] = 10^{-pOH} = 10^{-13}M$$

57. A 1M solution of HA, HB, HC and HD acids have pH values 1,2,3 and

4 respectively. Arrange these acids in increasing order of strength.

Solution

With same concentrations, an acid with lower pH value is stronger.

$\therefore$ HD < HC < HB < HA

58. A human blood has a pH of 7.4. Its pOH will be _____

Solution

pOH = 14 - pH = 14 - 7.4 = 6.6

59. What is the pH of 0.001M solution of KOH?

Solution

For strong mono base, [base] = $[OH^-]$

$\Rightarrow [KOH] = [OH^-] = 0.001M = 10^{-3}M$

$pH = 14 + 10g[OH^-] = 14 + 10g10^{-3} = 3$

60. Calculate the pOH of 0.1M $Ba(OH)_2$ solution

Solution

$Ba(OH)_2 \longrightarrow Ba^{2+} + 2OH^-$

1M 2M

0.01M ?

$$\Rightarrow [OH^-] = \frac{0.01M \times 2M}{1M} = 2\times10^{-2}M$$

$\therefore pOH = -10g[OH^-] = -10g \times 10^{-2} = 1.7$

61. What is the pOH of 0.02M H_2SO_4 solution?

Solution

$H_2SO_4 \longrightarrow 2H^+ + SO_4^{2-}$

1M 2M

0.02M ?

$$\Rightarrow [H^+] = \frac{0.02M \times 2M}{1M} = 0.04M = 4 \times 10^{-2}M$$

$pOH = 14 + 10g[H^+] = 14 + 10g4\times10^{-2} = 12.6$

Preparation of Acids

There are different methods used to prepare acids.

1. **The reaction between acidic oxides and water**

Acidic oxide + water → Acid

Example

$N_2O_3 + H_2O \rightarrow 2HNO_2$

$P_4O_6 + 6H_2O \rightarrow 4H_3PO_3$

2. **Direct combination of some non - metals with hydrogen.**

- This method is used to prepare binary acids

Example

- Hydrogen + chlorine → Hydrogen chloride

  $H_2(g) + Cl_2(g) \rightarrow 2HCl\ (g)$

- An aqueous solution of hydrogen chloride is known as hydro chloric acid (HCl (aq))
- Hydrogen + Sulphur → Hydrogen sulphide

  $H_2(g) + S(s) \rightarrow H_2S\ (g)$

- An aqueous solution of hydrogen sulphide is called hydro sulphuric acid (H_2S (aq))

3. Using **double displacement reactions**

- Volatile acids can be prepared by reacting their salts with non - volatile acids.

Example

Sodium chloride + Sulphuric acid → sodium hydrogen sulphate + Hydrogen acid

$NaCl(s) + H_2SO_4(1) \rightarrow NaHSO_4(s) + HNO_3(1)$

Note: HCl and HNO_3 are volatile acid.

H_2SO_4 is non - volatile acid.

Preparation of Bases

Bases can be prepared by:

1. **The reaction between basic oxide and water**

Base oxide + water → Base

Example:

$Na_2O + H_2O \rightarrow 2NaOH$

$CaO + H_2O \rightarrow Ca(OH)_2$

2. **The reaction between active metals and water**

Active metal + water → Base + hydrogen

Example

$2Na + 2H_2O \rightarrow 2NaOH + H_2$

$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$

3. **Double displacement reaction**

☑ The reaction between soluble base and its salt results another soluble base and insoluble salt.

Soluble base + soluble salt → soluble base + Insoluble salt

Example

$Ba(OH)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaOH(aq)$

$Ca(OH)_2(aq) + K_2CO_3(aq) \rightarrow 2KOH(aq) + CaCO_3(s)$

Note: Industrially, KOH and NaOH are manufacture from electrolysis of concentrated solution of KCl and NaCl respectively.

Illustrative Examples

62. Which type of acids is prepared by direct combination of non - metal with hydrogen?

A) Ternary acids B) Organic acids C) Binary acids D) Oxy acids

Solution

Non - Metal + Hydrogen → Binary acid

Answer: C

63. Which of the following is volatile inorganic acid?

A) HCl and HNO_3 C) HCOOH and CH_3COOH

B) H_2SO_4 and HNO_3 D) HCOOH and HCl

Answer: A

64. Given the reaction: $P_4O_{10} + H_2O \rightarrow X$ The chemical name of X is:

A) Phosphorus acid C) Hypophosphorus

B) Phospheric acid D) Phosphine

Solution

$P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$ (Phospheric acid)

H_3PO_3 - Phosphorus acid PH_3 - Phosphine

H_3PO_2 - Hypophosphorus acid

Answer: B

65. Which one of the following processes does not form a hydroxide base?

A) Reaction of active metal with water

B) Reaction of alkali with a solution of salt

C) Direct combination of elements

D) Reaction of soluble metallic oxide with water

Solution

Active metal + water → Hydroxide base

Alkali (soluble base) + water soluble → Hydroxide base + insoluble salt

Metallic oxide + water → Hydroxide base

Answer: C

66. Which of the following reactions results a base?

A) Reaction of acidic oxide with water

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B) Reaction of active metals with water

C) Reaction of metals with hydrogen

D) Reaction of acids with bases

67. Complete and balance the following

Answer: B

a) $SO_3 + H_2O \rightarrow$

b) $H_2 + F_2 \rightarrow$

c) $H2SO4 + KCl \rightarrow$

d) $K + H_2O \rightarrow$

e) $BaO + H_2O \rightarrow$

f) $Ca(OH)_2 + NaCO_3 \rightarrow$

Solution

a) $SO_3 + H_2O \rightarrow H_2SO_4$

b) $H_2 + F_2 \rightarrow 2HF$

c) $H_2SO_4 + KCl \rightarrow KHSO_4 + HCl$

d) $2K + 2H_2O \rightarrow 2KOH + H_2$

e) $BaO + H_2O \rightarrow Ba(OH)_2$

f) $Ca(OH) + Na_2CO_3 \rightarrow CaCO_3 + 2NaOH$

68. Define alkalis and give example

Solution

Alkalis are soluble bases

Example: NH_3, LiOH, NaOH, KOH, RbOH, $Ca(OH)_2$, $Ba(OH)_2$, $Sr(OH)_2$

Some important Acids and Bases

1. Sulphuric acid (H_2SO_4)

It is rank number one important chemical that determines the development of a country.

The major uses of sulphuric acid are:

- To manufacture fertilifers, dyes, paints, explosive, drugs, rayon (synthetic fibers), detergents.
- Used to clean metal surfaces.
- Used as electrolyte in car batteries

- Used as oxidizing and dehydrating agents
- Used as catalyst in organic chemistry

For example, to make esters from alcohol and carboxylic acid by esterification reaction.

2. Hydrochloric Acid (HCl)

It is used for:

- Cleaning metal surface before electroplating
- Preparation of drugs, dyes and photographic chemicals
- Manufacturing chlorine gas

3. Nitric acid (HNO_3)

- Pure nitric acid is known as "aqua fortis"
- It is used to make fertilizers, explosives such as trinitrtoluene (TNT) and trinitro glycerine.
- It is used to manufacture rubber, chemicals, plastics, dyes and drugs.

4. Phospheric acid (H_3PO_4)

It is used as:

- Rust converter or remover for metal surfaces
- Carbonator and flavoring agent for beverages
- An etching solution for dentistry and semiconductor processing
- A fertilizer ingredient for making phosphate salts
- A component of dental cements, albumin derives, sugar and textile.

5. Acetic Acid (CH_3COOH)

Acetic acid has many uses. Some of them are

- Industrial use, medicinal use, house hold use and food industry.

6. Sodium Hydroxide (NaOH)

It is commonly known as **Caustic soda or lye** because it is corrosive

It is used in:

- Dry cells as electrolytes
- Making soap, pulp, paper and textile
- Purifying aluminum ores and removing sulphur from petroleum.

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7. Potassium Hydroxide (KOH)

It is also called **caustic potash**

It is stronger than sodium hydroxide

➥ KOH is used to make soft soaps and as electrolyte in alkaline batteries.

8. Calcium Hydroxide ($Ca(OH)_2$)

It is also known as **slacked lime or lime water.**

➥ It is widely used to remove the acidity of soil

➥ It is used to remove temporarly hardness of water

➥ It is used to remove hairs from hides in tanning process.

➥ It is also used as a component of **mortar** (mixture of slacked lime, sand and water)

9. Ammonia (NH_3)

A solution of ammonia in water forms ammonium hydroxide (NH_4OH)

It is used as:

➥ Household cleaning agent

➥ Laboratory reagent.

➥ Raw material to make nitric acid, sodium carbonate and fertilizers such as ammoniunitrate

10. Magnesium Hydroxide ($Mg(OH)_2$)

It is also known as **milk of magnesia**

Some of its uses are

➥ As antiacid or a laxative in oral liquid or chewable tablet form

➥ As a thermal conductor and poor electrical conductor

➥ As a food additive

➥ As a waste - water treatment agent.

➥ As a fire retardent and smoke suppressor.

Illustrative Examples

69. Which of the following is used as a raw material in the production of nitric acid?

A) $NaNO_3$ B) N_2O C) N_2O_3 D) NH_3

Solution

Ammonia is used to make nitric acid by Ostwald process.

$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

$2NO + O_2 \rightarrow 2NO_2$

$3NO_2 + H_2O \rightarrow 2HNO_3 + NO$

Answer: D

70. Which of the following is not true about ammonia?

A) Ammonia is used to make fertilizers

B) Ammonia is used as a cleaning agent

C) Ammonia is used as anesthetic in dental and minor surgery

D) Ammonia is used as refrigerant

Answer: C

71. A country is said to be developed by its large quantity production and consumption of:

A) Sulphuric acid C) Phospheric acid

B) Nitric acid D) Hydrochloric acid

Solution

The development of a country can be determined by the amount of sulphuric acid it produces and consumes.

Answer: A

72. One of the most important industrial application of nitric acid is

A) For dissolving some metals that cannot be dissolved by other acids.

B) For production of photographic films

C) For an oxidizing agent

D) For manufacturing of fertilizers

Answer: D

73. Write the formula and name of two important feritilizers of nitric acids

Solution

➥ NH_4NO_3 - Ammonium nitrate

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➥ KNO_3 - Potassium nitrate

74. A base commonly known as slacked limed and used to treat acidity of soil is _____.

Answer: Calaium hydroxide ($Ca(OH)_2$)

75. List the industrial applications of acetic acid or ethanoic acid.

Solution

Chemical reagent, fungicide, herbicide, microbiocide and pH adjuster.

Precautions while Working with Acids and Bases in Laboratory

- ☑ It is important to wear goggles, gloves and laboratory coat when dealing with acids and bases.
- ☑ Use bellows to pipet acidic or basic solutions rather than sucking it with your mouth.
- ☑ If concentrated acid or base is spilled on cloth immediatery wash with water
- ☑ If acid enters your eye, wash it with water then consult doctor.
- ☑ If acid accidentally swallowed, administer weak bases such as magnesium hydroxide ($Mg(OH)_2$) or aluminum hydroxide ($Al(OH)_3$).
- ☑ If a base is swallowed by accident, drink 1 - 2% acetic acid or lemon juice.
- ☑ Whenever bases come into contact with your skin, wash the affected part with water and then by weak acid solution such as acetic acid.
- ☑ During dilution of concentrated acid, add the acid in to water (AW) but not water to acid (WA). Follow the Amharic **AW!** Not **WA!** (A - Acid, W - Water)

Illustrative Examples

76. Which of the following is safe to perform in the laboratory?

A) Administering strong acid for swallowed base.

B) Adding concentrated acid carefully to water in order to dilute

C) Washing an acid entered the eye with water then adding a weak base

D) Washing a splash concentrated acid on the body with water then treating with NaOH

Answer: B

77. Which of the following safety rules are not allowed in handling bases in high school laboratories?

A) Wearing eye goggles, gloves and laboratory coat

B) Washing the affected part with water whenever bases are in contact with our skin

C) Cleaning the working table immediatery if bases are spilled on it

D) Tasting bases using our tangue.

Solution

Tasting chemicals is not allowed in laboratories even chewing gum and drinking water.

Answer: D

78. List additional safety rules.

Solution

- ➥ Do not point reaction vessels directly toward you and your friends.
- ➥ Wear shoes that cover you feet
- ➥ For volatile substances, use gashood.
- ➥ Keep your hands away from your face, eyes, mouth and body while used chemicals
- ➥ Never use laboratory glass ware for eating or drinking purposes.

8.3 Salts

Salts are ionic compounds obtained by a neutralization reaction.

- ➥ They are formed when a metal or ammonium ion takes the place of hydrogen (s) in an acid
- ➥ **The positive ion (Cation)** of a salt is obtained from base and its **negative ion (anion)** from acid.

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Example

$$\overset{+}{K}\overset{-}{OH} + \overset{+}{H}\overset{-}{I} \longrightarrow \overset{+}{K}\overset{-}{I} + H_2O$$

Base Acid Salt

Table 8.4 Some Acids and their Salts

Acid	Its salt
Sulphuric acid	Sulphate
Sulphurous acid	Sulphite
Nitric acid	Nitrate
Nitrous acid	Nitrite
Phospheric acid	Phosphate
Phosphorus acid	Phosphite
Perchloric acid	Perchlorate
Chloric acid	Chlorate
Hydrochloric acid	Chloride
Hydrosulphic acid	Sulphide
Hydrocyanic acid	Cyanide

Classification of Salts

Salts can be classified into three groups. These are

1. Normal salt
2. Acidic salt
3. Basic salt

1. Normal salts /Neutral salts/

These are salts in which all the ionizable hydrogen of acids are replaced metal or ammonium ions.

Example

$NaCl$, KNO_3, CH_3COONH_4, $CaSO_4$, BaI_2 etc

Note

i. **A normal salt** obtained from weak acid and strong base is **basic salt.** ($pH > 7$)

Example:

Na_2CO_3, CH_3COONa, KF, CaS etc

ii. **A normal salt** formed from strong acid and weak base is **acidic salt** ($pH < 7$)

Example

NH_4Cl, $MgSO_4$, $Al(NO_3)_3$ etc

iii. **A normal salt** formed from strong acid and strong base **neutral salt.** ($pH = 7$)

Example

$NaCl$, K_2SO_4, CaI_2, $Ba(NO_3)_2$ etc

iv. **A normal salt** obtained from weak acid and weak base can be **neutral, acidic or basic.**

2. Acidic Salts

These are salts that contain one or more ionizable hydrogens in their formula.

Example

$NaHCO_3$ - Sodium hydrogen carbonate

$KHSO_4$ - Potassium hydrogen sulphate

$Ca(H_2PO_4)_2$ - Calcium dihydrogen phosphate

$CaHPO_4$ - Calcium monohydrogen phosphate etc.

3. Basic Salts

These are salts that have one or more hydroxyl groups.

Example

$Pb(OH)Cl$ - Basic lead (11) chloride

$Cu(OH)NO_3$ - Basic copper (11) nitrate

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$Zn(OH)NO_3$ - Basic zinc nitrate etc.

Based on the type of element they contain salts can be classed as

1. **Binary salts** are salts formed from two types of elements.

Example

$AlCl_3$ - Aluminum chloride

KI - Potassium iodide

$CaBr_2$ - Calcium bromide etc.

2. **Ternary salts** are formed from three types of elements.

Example

$Al_2(SO_4)_3$ - Aluminum sulphate

KNO_3 - Potassium nitrate

$NaClO_4$ - Sodium perchlorate etc.

Illustrative Example

79. A ternary salt among the following is:

A) Al_2S_3 C) KH_2PO_4

B) CH_3COONa D) $Ca(ClO_4)_2$

Solution

Ternary salts are composed of three types of elements

Answer: D

80. One of the following salts is acidic salt. Which one is it?

A) $Ca(NO_3)_2$ C) KH_2PO_4

B) $Zn(OH)Cl$ D) Na_2SO_4

Solution

Acidic salts contain one or more ionizable hydrogens in their formula.

Answer: C

81. Identify the pair of compounds that is not related like the other pairs.

A) Nitric acid - sodium nitrate

B) Sulphuric acid - sulphurdioxide

C) Carbonic acid - calcium carbonate

D) Phospheric acid - potassium phosphate

Answer: B

82. Classify the following salts as normal, acidic or basic salts.

a) $LiHCO_3$ c) KHS e) NH_4Cl

b) $BaSO_4$ d) PB (OH)Cl f) $CuCl_2$

Solution

- a and c are acidic salts
- b, e and f are normal salts
- d is basic salt

83. Categorize the following salts as binary or ternary salts

a) MgF_2 b) $FeSO_4$ c) KCN d) CuF_2

Solution

- a and d are binary salts
- b and c are ternary salts

84. Write the name of the following salts

a) FeF_3 c) Li_3N e) $Zn(OH)NO_3$

b) $CuSO_4$ d) $Ca_3(PO_4)_2$ f) $NaHSO_3$

Solution

a) Iron (III) fluoride

b) Copper (II) sulphate

c) Lithium nitride

d) Calcium phosphate

e) Basic zinc nitrate

f) Sodium hydrogen sulphite

85. Why salts are called ionic compounds?

Solution

Because they are aggregates of positive and negative ions.

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Preparation of Salts

Five Methods for Preparing Salts are as Follows;

1. Neutralization Reaction

An acid and a base neutralize each other to form the appptopriate salt and water

Acid + Base → Salt + water

Example

$2HCl\ (aq) + Ca(OH)_2\ (aq) \rightarrow CaCl_2\ (aq) + H_2O(1)$

This method is used to prepare, soluble salt, insoluble salt, binary and ternary salt.

$NaOH\ (aq) + HCl(aq) \rightarrow NaCl\ (aq) + H_2O(1)$ soluble salt

$Ba(OH)_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + H_2O(1)$ Insoluble salt

$KOH(aq) + HBr(aq) \rightarrow KBr(aq) + H_2O(1)$ Binary salt

$Sr(OH)_2(aq) + 2HNO_3(aq) \rightarrow Sr(NO_3)_2\ (aq) + 2H_2O(1)$ Ternary salt

2. Single Replacement Reaction

An active metal replaces hydrogen in an acid

Active metal + Acid → Salt + Hydrogen

Example

$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

3. Direct Combination of Elements

It is the combination of active metals with non - metals, such as halogens, sulphur, phosphorus and nitrogen to yield binary salt.

Metal + non - metal → salt

Example

$Fe(s) + S(s) \rightarrow FeS(s)$

4. Double Displacement Reaction

When solutions of two soluble salt are mixed, they form soluble and in

soluble salts (Precipitation)

Soluble salt + Soluble salt → Insoluble salt + soluble salt

Example

$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl\ (s) + NaNO_3(aq)$

5. Reaction of Acidic oxide with Basic Oxide

Acidic oxide + Basic oxide → salt

Example

$MgO\ (s) + SiO_2\ (s) \rightarrow MgSiO_3\ (s)$ (magnesium silicate)

Note

i. Acid + carbonate/hydrogen carbonate → salt + carbon dioxide + wate

ii. Acid + basic oxide → salt + water

iii. Base + acidic oxide → salt + water

Illustrative Examples

86. Salts can be prepared by the reaction between

A) Bases and non - metals

B) Metallic oxide and dilute acids

C) Acidic oxide and water

D) Active metal and water

Solution

Base + non - metal → No reaction

Metallic oxide + acid → salt + water

Acidic oxide + water → Acid

Active metal + water → Base + hydrogen

Answer: B

87. What would be the formula of the salt that might be produced when carbonic acid (H_2CO_3) is combined with aluminum hydroxide ($Al(OH)_3$)?

A) $AlCO_3$ B) Al_4C_3 C) $Al_2(CO_3)_3$ D) Al_2CO_3

Solution

$Al(OH)_3 + H_2CO_3 \rightarrow Al(HCO_3)_3 + H_2O$

partial replacement of hydrogen

$2Al(OH)_3 + 3H_2CO_3 \rightarrow Al_2(CO_3)_3 + 6H_2O$

complete replacement of hydrogen

Answer: C

88. Which of the following acids forms three types of salts?

A) Carbonic acid C) Phospheric acid

B) Perchloric acid D) Sulphuric acid

Solution

Monoprotic acids form one type of salt. Diprotic acids result two types of salts. Triprotic acids produce three types of salts, H_3PO_4 is a triprotic acid that forms three types of salts. For example NaH_2PO_4, $NaHPO_4$ and Na_3PO_4.

Answer: C

89. Identify the acid and the base used to prepare:

a) $NaNO_3$ c) KCN e) $CuCO_3$

b) $Al_2(SO_4)_3$ d) BaS f) $Fe(ClO_4)_3$

Solution

Acid	Base
a) HNO_3	NaOH
b) H_2SO_4	$Al(OH)_3$
c) HCN	KOH
d) H_2S	$Ba(OH)_2$
e) H_2CO_3	$Cu(OH)_2$
f) $HClO_4$	$Fe(OH)_3$

90. Write the formula and name of salts formed when potassium hydroxide react with sulphuric acid.

Solution

$KHSO_4$ - Potassium hydrogen sulphate.

K_2SO_4 - Potassium sulphate

91. Complete and balance the following

a) $Zn + HI \rightarrow$

b) $CuO + H_2SO_4 \rightarrow$

c) $KHCO_3 + HNO_2 \rightarrow$

d) $CaCl_2 + Na_2CO_3 \rightarrow$

Solution

a) $Zn + 2HI \rightarrow ZnI_2 + H_2$

b) $CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$

c) $KHCO_3 + HNO_2 \rightarrow KNO_2 + H_2O + CO_2$

d) $CaCl_2 + Na_2CO_3 \rightarrow CaCO_3 + 2NaCl$

92. How could you prepare insoluble salts?

Solution

By using double replacement reaction solutions of two soluble salt result one soluble and one insoluble salt. The insoluble salt precipitates or solidifies and separated easily by filtration

93. How can you prepare silver acetate by using neutralization reaction?

Solution

$2CH_3COOH + Ag_2O \rightarrow 2CH_3COOAg + H_2O$

$CH_3COOH + AgOH \rightarrow CH_3COOAg + H_2O$

Some Important Salts

Table 8.5 Some Important Salts and their Uses

Salt	Used
NaCl - sodium chloride - table salt	➥ It is for flavoring and preserving food ➥ It is used as a raw material to produce Na, Cl_2 NaOH, $NaHCO_3$ and Na_2CO_3. ➥ It is used in tannery to dry the skin and kill bacteria ➥ It is component of oral rehydration salt (ORS)

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$CuSO_4$ Copper (II) sulphate (cupric sulphate)	➥ It is useful in electroplating ➥ It is used to prevent fungal attack of leaves and vines ➥ It is used to make a mixture of $CuSO_4$ and $Ca(OH)_2$ called **Bordeaux**
$FeCl_3$- Iron (III) chloride (Ferric chloride	➥ It is used to treate waste water and feed additive ➥ It is used to etch printed circuit
$FeSO_4$ - Iron (II) sulphate (Ferric sulphate	➥ It is used to iron tablets to protect patients from anaemia. ➥ It is used as wood preservative, water and sewage treatment ➥ It is used to manufacture ink
$BaSO_4$- Barium sulphate	➥ It is used as white pigment and taking X - ray photographs of the intestinal tract since it is opaque to X - rays
$CaSO_4 . 2H_2O$ - hydrated calcium sulphate (gypsum)	➥ It is used to plaster wall and support fracture bones
KNO_3 - Potassium nitrate (Salt peter)	➥ It is used to make fertilizer ➥ It is component of gun powder ($KNO_3 + C + S$)
NH_4NO_3 - ammonium nitrate	➥ It is mainly used as nitrogeneous fertilizer
$Al_2(SO_4)_3$ - Aluminum sulphate	➥ It is mainly used in paper industry and water purification

$MgSO_4 .4H_2O$ - Hydrated magnesium sulphate (Epsom salt)	➥ It is used as laxative and for other medical purpose

Illustrative Example

94. Which of the following salts is used to make fungicides?

A) $CuSO_4$ B) NaCl C) KNO_3 D) $BaSO_4$

Answer: A

95. Identify the salt that is given as iron tablets to patients who suffer from anaemia among the following?

A) NaCl B) NH_4NO_3 C) $FeSO_4$ D) $FeCl_3$

Answer: C

96. Which of the following salts supports fractured bones?

A) NH_4NO_3 C) $MgSO_4 .4H_2O$

B) $CaSO_4 - 2H_2O$ D) $FeCl_3$

Answer: B

97. Write the components of

a) Gun powder b) Bordeaux

Solution

a) KNO_3, S and C b) $CuSO_4$ and $Ca(OH)_2$

98. List the substance formed by using NaCl

Solution

Na, Cl_2, NaOH, $NaHCO_3$ and $NaCO_3$

99. Write the formula of epsom salt

Answer: $MgSO_4.7H_2O$

Properties of Salts

1. Solubility of Salts in Water

☑ Salts containing the following ions are soluble in water.

- Alkali metals (Li^+, Na^+, K^+, Rb^+)
- Ammonium ion (NH^+_4), nitrate ion (NO^-_3), perchlorate ion (ClO^-_4), chlorate ion (ClO^-_3)
- Acetate (CH_3COO^-)

☑ All halides except that of Ag^+, Hg^{2+} and Pb^{2+} are soluble in water

☑ All sulphates except that of Ba^{2+}, Hg^{2+}, Pb^{2+} and Sr^{2+} are soluble in water

☑ Carbonates, phosphates and sulphides except that of alkali metals and ammonium ion are insoluble in water.

☑ Sulphides of group IIA metals are soluble in water.

2. Some Salts Absorb and other Release Water

Salts that absorb water from the atmosphere but remain solid are called **hygroscopics salts.**

Example

$CuSO_4$ - Anhydrous copper (II) sulphate

Salts that absorb water from the atmosphere and form solution are called **deliquescent salts**

Example

$CaCl_2$ - Calcium chloride

$NaNO_3$ - Sodium nitrate

$FeCl_3$ - Iron (III) chloride

$MgCl_3$ - Magnesium chloride

$LiNO_3$ - Lithium nitrate

Salts that lose water of crystallifation to the atmosphere are known as **efflorescent salts**

Example

$Na_2CO_3 . 10H_2O$ - Hydrated sodium carbonate

$Na_2SO_4 . 10H_2O$ - Hydrate sodium sulphate

Note:
All deliquescent substance are hygroscopic, but all hygroscopic sub stances are not necessarily deliquescent.

3. Salts are Electrolytes

Molten and aqueous states of salts transmit electricity due to the mobile negative and positive ions.

4. Salts are Decomposed by Heat

i. Thermal Decomposition of Carbonates

- Carbonates are salts of carbonic acid
- They decomped thermally and give metallic oxide and carbondioxide.

Example

$CaCO_3 \rightarrow CaO + CO_2$

$Li_2CO_3 \rightarrow Li_2O + CO_2$

Na_2CO_3 and K_2CO_3 are not decomposed by heat

ii. Thermal Decomposition of Nitrates

- Nitrates are salts of nitric acid
- They give metallic oxide, nitrogen dioxide and oxygen up on decomposition

Example

$2Pd(NO_3)_2 \rightarrow 2PbO + 4NO_2 + O_2$

$2Li\,NO_3 \rightarrow 2Li_2O + 4NO_2 + O_2$

$NaNO_3$ and KNO_3 give metal nitrite and oxygen upon reduction by heat

$2NaNO_3 \rightarrow 2NaNO_2 + O_2$

$2KNO_3 \rightarrow 2KNO_2 + O_2$

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Illustrative Examples

100. All of the following salts are soluble in water. **EXCEPT**

A) NH_4NO_3 B) KCl C) Na_2CO_3 D) $BaSO_4$

Answer: D

101. Which of the following salts cannot decomposed by heat?

A) K_2CO_3 B) $KClO_3$ C) KNO_3 D) None

Solution

Na_2CO_3 and K_2CO_3 are not decomposed by heat

Answer: A

102. Which of the following is deliquescent salt?

A) $Na_2SO_4 . 10H_2O$ C) $CuSO_4$

B) $NaNO_3$ D) $Na_2CO_3.10H_2O$

Answer: B

103. Complete and balance the following

a) $NaNO_3 \xrightarrow{\Delta}$ c) $Ca(NO_3)_3 \xrightarrow{\Delta}$

b) $LiNO_3 \xrightarrow{\Delta}$ d) $MgCO_3 \xrightarrow{\Delta}$

Solution

a) $2NaNO_3 \xrightarrow{\Delta} 2NaNO_2 + O_2$

b) $4LiNO_3 \xrightarrow{\Delta} 2Li_2O + O_2 + 4NO_2$

c) $2Ca(NO_3)_2 \xrightarrow{\Delta} 2CaO + 4NO_2 + O_2$

d) $MgCO_3 \xrightarrow{\Delta} MgO + CO_2$

104. What types of salts release a redish - brown gas nitrogen dioxide when they are decomposed by heat?

Answer: Metal nitrates except that of sodium and potassium

105. The process of absorbing water from the atmosphere by a solid to form a solution is knownas _____

Answer: Deliquescent

8.4 Chemical Tests of Some Ions of Salts

1. Flame Test

- It is a qualitative test to identify metal ions in salts
- The type of alloy used for flame test is called **nichrome wire** (an alloy of nickel and chromium)

Table 8.6 Flame Test Result of Some Cations

Metal ion in salt	Color of flame produced
Lithium (Li^+)	Crimson
Sodium (Na^+)	Yellow
Potassium (K^+)	Violet
Calcium (Ca^{2+})	Orange red
Strontium (Sr^{2+})	Crimson
Barium (Ba^{2+})	Green

Blue solution of copper indicates the present of Cu^{2+} in the solution.

Pale -green precipitate formation when iron (II) compounds react with alkaline solution indicates presence of Fe^{2+}

Pale - brown precipate formation when iron compounds react with alkaline solution confirms the presence of Fe^{3+}

2. Test of Halides

Mixing sodium halides with silver nitrate, $AgNO_3$

When **white precipitate** is formed, it confirms the presence of **chloride ion** (Cl^-)

Formation of **yellow precipitate** indicates the presence of **bromide ion** (Br^-)

Formation of **yellowish - green** precipitate proves the presence of **iodide ion** (I^-)

3. Test for Sulphates (SO_4^{2-})

When soluble salts of barium react with sodium salts, **White precipitate,**

$BaSO_4$ is formed. This confirms the presence of sulphate.

4. Test for Nitrates (NO^-_3)

Nitrates are identified by **Brown ring test**

In a freshly prepared iron (II) sulphate, $FeSO_4$ if concentrated sulphuric acid is poured down the side of the test tube, the acid sinks to the bottom and a brown ring is formed. This happens when there is nitrate ion in the solution

The formula of brown ring is **$FeSO_4$. NO**

Illustrative Examples

106. Which of the following group of salts give a white precipitate when a solution of barium chloride is added to them?

A) Nitrates C) Permanganate

B) Chlorates D) Sulphates

Answer: D

107. Which of the following could form a black precipitate when lead acetate is added to it?

A) H_2SO_4 B) H_2S C) SO_2 D) SO_3

Solution

$$Pb(CH_3COO)_2(aq) + H_2S(aq) \rightarrow PbS(s) + 2CH_3COOH(aq)$$

black precipitate

Answer: B

108. The Brown ring test is used to identify the presence of:

A) NO^-_3 B) SO^{2-}_4 C) NO^-_2 D) SO^{2-}_3

Answer: A

109. What is the laboratory apparatus used for flame test?

Answer: Nichrome wire

110. The reaction between sodium salt and silver nitrate gives white precipitate. This confirms the presence of _____

Answer: Chloride ion (Cl)

8.5 Plant Nutrients

Plant nutrients are minerals required for the growth and development of plants. These nutrients are classified in to two

Plant Nutrients

Macro Nutrients
→ required for plants in larger quantity

Example
- Nitrogen
- Phosphorus
- Potassium
- Magnesium
- Calcium
- Sulphur
- Carbon
- Hydrogen
- Oxygen

Micro Nutrients
→ required for plants in smaller quantity

Example
- Iron
- Cobalt
- Zinc
- Manganese
- Molybdenum
- Copper
- Boron
- Chlorine

The role of nitrogen, phosphorus and Potassium in plants growth

A. Nitrogen

- ☑ It is absorbed by plants in the form of nitrate ion.
- ☑ It is used to make amino acids, proteins, coenzymes, and nucliec acids.
- ☑ It is also used in the synthesis of the green pigment chlorophyll
- ☑ The leaves of plants turn yellow when they face lack of nitrogen.

B. Phosphorus

- ☑ It is absorbed by plants in the form of dihydrogen phosphate ($H_2PO^-_4$)
- ☑ It facilitates early growth. root formation, quick maturity and promotes

seed or fruit production.

☑ It has role in the formation of amino acids, protein, coenzmes, nucleic acids and high energy phospate (ATP)

C. Potassium

☑ It is absorbed by plants in the form of potassium ion (K^+)

☑ It facilitates the process of photosynthesis

☑ It adjusts stomata movements.

☑ It controls the turgidity of plants

☑ It also increases solute potential

Fertilizers

☑ **Fertilizers** are added to soil, to increase the growth, yield and nutrient value of crop.

☑ They are also used to replace the nutrients lost from the soil. There are two types of fertilizers.

Table 8.7 The Two Type of Fertilizers

Organic/natural fertilizers	Synthetic /Artificial/fertilizers
➥ They are obtained from animals and plants ➥ These fertilizers can supply the elements nitrogen, phosphorus and varius trace elements to the soil **Example** Animal dung, urine and substance from decay of plants	➥ Contain all the three major plants mutrients K,P and N ➥ They are complete fertilizers **Example** ➥ Diammonium monohydrogen phosphate, DAP, $(NH_4)_2 HPO_4$ ➥ Ammonium sulphate, $(NH_4)_2 SO_4$ ➥ Potassium nitrate (KNO_3)

Pesticides

These are chemicals that kill pests which affect the growth of plants, most pesticides are organic compounds.

They are used to over come the damage caused by fungi and pests.

Example

➥ Fumigants: HCN, SO_2, CS_2 and PH_3

➥ Inorganic pesticides: $CuCl_2$, $Cu(OH)_2$, $Cu(NO_3)_2$, $CuSO_4$, Cu_2O and $CuCO_3$

➥ $CuSO_4$ and Cu_2O are fungicides.

➥ HCN, SO_2, PH_3, $CuCl_2$, $Cu(OH)_2$ and $CuCO_3$ are used as insecticides

Herbicides /Weed killers/

➥ They are chemicals that remove weeds

➥ Weeds are unwanted plants

Example

Borax : $Na_2 B_4O_7 . 10H_2O$

Sodium chlorate: $NaClO_3$

Hydrated iron (II) Sulphate: $FeSO_4 . 7H_2O$

Ammonium sulphamate: $(NH_4) SO_3 . NH_2$

Illustrative Examples

111. Which of the following set elements are macro nutrients of plants?

A) Zn, Fe, B B) Cl, Cu, Mo C) Fe, Cl, Mn D) P, K, N

Answer: D

112. Which of the following is a pesticide?

A) Borax C) Potassium nitrate

B) Copper (II) sulphate D) Sodium chlorate

Solution

➥ Sodium chlorate and borax are herbicides

➥ Potassium nitrate is a fertilizer.

Answer: B

113. What type fertilizer is a complete fertilizer?

Solution

A fertilizer that contains the three major elements potassium,

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phosphorus and nitrogen is called complete fertilizer.

114. Write the synthetic fertilizer that contains two of the major macro elements

Solution

DAP, $(NH_4)_2 HPO_4$

Solved problems

115. Inorganic compounds are originated from _____.

Solution

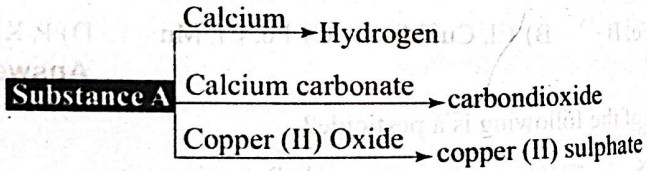
Mineral constituents of the earth's crust.

116. An unknown substance reacts with marble chips releasing a gas that estinguishes with marble chips and neutralizes a basic solution. The unknown substance could most likely be _____.

Solution

A gas that extinguishes fire is carbon dioxide. Thus, the unknown substance that releases CO_2 when it reacts with marble chips is an acid.

117. Some reaction of a substance, A, are shown in the diagram below.



Answer: Acid

118. What volume of 0.5M HCl contains 10moles of HCl?

Solution

$$V = \frac{n}{M} = \frac{10mol}{0.5\frac{mol}{L}} = 20L$$

119. The pH of 250ml of H_2SO_4 solution is one. Calculate the mass of H_2SO_4 required to prepare this solution

Solution

$[H^+] = 10^{-PH} = 10^{-1} M = 0.1M$

$H_2SO_4 \longrightarrow 2H^+ + SO^{2-}4$

1M 2M

? 0.1M

$$\Rightarrow [H_2SO_4] = \frac{1M \times 0.1M}{2M} = 0.05M$$

$M_{wt} = 98 g/mol$

$$m = M \times M_{wt} \times V = 0.05\frac{Møl}{L} \times 98\frac{g}{møl} \times 0.25L$$

$m = 1.225g$

120. When sodium hydroxide is added to dilute hydrochloric acid solution in a beaker. What happens to the pH, pOH, $[H^+]$ and $[OH^-]$ values?

Solution

pH and $[OH^-]$ increase

pOH and $[H^+]$ decrease

121. Two solutions A and B have pH values 3 and 6. How many times greater is the hydrogen ion concentration in solution A that of solution B?

Solution

$$\frac{[H^+]_A}{[H^+]_B} = \frac{10^{-pHA}}{10^{-pHB}} = \frac{10^{-3}}{10^{-6}} = 1000$$

$$\Rightarrow [H^+]_A = 1000[H^+]_B$$

122. Calculate the pH of 1×10^{-3}M Ba $(OH)_2$ solution.

Solution

$[OH^-] = 2 \times 1 \times 10^{-3} M = 2 \times 10^{-3} M$

$pH = 14 + 10g[OH^-] = 14 + 10g\,2 \times 10^{-3} = 11.30$

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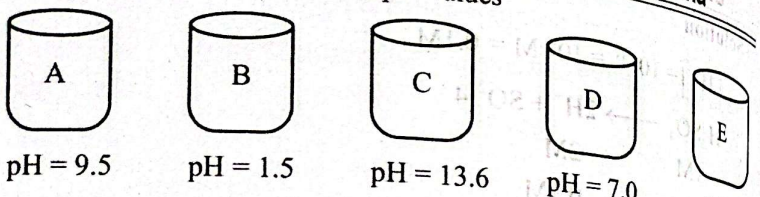
123. Given solution with different pH Values



Which solution is

a) Neutral?
b) Slightly acidic?
c) Strongly basic?
d) Strongly acidic?
e) Slightly basic?

Solution

a) D b) E c) C d) B e) A

124. Ammonia is a weak base:

a) What do you understand by the term weak base?
b) Write a balanced equation to show the reaction of ammonia with sulphuric acid.

Solution

a) It is a base that ionizes slightly when dissolved in water.
b) $2NH_3 + H_2SO_4 \rightarrow (NH_4)_2 SO_4$

125. Explain the amphoteric character of zinc oxide by using HNO_3 and NaOH.

Solution

$ZnO\ (s) + 2HNO_3\ (aq) \rightarrow Zn\ (NO_3)_2\ (aq) + H_2O\ (l)$
Act as base

$ZnO(s) + 2NaOH\ (aq) \rightarrow Na_2\ ZnO_2\ (aq) + H_2O(l)$
Act as an acid Sodium zincate

126. _____ is a measure of the concentration of the hydrogen ions (H) in a solution or it is the acidity or alkalinity of that solution.

Answer: pH

127. Calculate the pH of as solution formed by dissolving 2.8g of KOH in 100ml of solution.

Solution

$$[KOH] = [OH^-] = \frac{m}{M_{wt} \times V} = \frac{2.8g}{56\frac{g}{mol} \times 0.1L} = 0.5M$$

$$pH = 14 + 10g\,[OH^-] = 14 + 10g\ 0.5 = 13.70$$

128. State the name of a suitable acid and alkali you can use to make each of the following salts

a) Sodium sulphate
b) Potassium phosphate
c) Ammonium nitrate

Solution

Acid	Base
a) Sulphuric acid	Sodium hydroxide
b) Phospheric acid	Potassium hydroxide
c) Nitric acid	Ammonia

129. Arrange the following acids in decreasing order of acid strength $HClO_3$, HClO, $HClO_2$ and $HClO_4$

Solution

$HClO_4 > HClO_3 > HClO_2 > HClO$ because as number of oxygen atom increases acidity also increases.

Supplementary Exercise

I. True/False items

1. Strong acids and strong bases ionize almost completely in aqueous solution.
2. Strongly basic solution has lower pH value than weakly basis solution.
3. pOH is defined as the logarithm of hydroxide ion concentration.
4. A solution is said to have basic character when its pH value is greater than seven.
5. When acidic substance is added to a given solution, the pH of the solution increase.
6. When an acid reacts with active metal, hydrogen gas is produced

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7. Iron (II) sulphate and sulphuric acid are reagents used to test nitrate ion.
8. A halide ion is tested with silver nitrate.
9. Farmers often add calcium oxide to treat the alkalinity of their soil.
10. Organic acids are generally stronger than inorganic acids

II. Matching items

A	B
11. KOH	A) Slacked lime
12. NaOH	B) Chile salt peter
13. $Ca(OH)_2$	C) Table salt
14. CaO	D) Caustic soda
15. $NaNO_3$	E) Caustic potash
16. KNO_3	F) Epsom salt
17. $MgSO_4 . 2H_2O$	G) Milk of magnesia
18. $CaSO_4 . 2H_2O$	H) Quick lime
19. NaCl	I) Gypsum
20. $Mg(OH)_2$	J) Salt peter
	K) Baking soda
	L) Washing soda

III. Multiple Choice items

21. The difference between HCl and CH_3COOH as acids is
   A) The first has less hydrogen in solution
   B) The first is highly ionized
   C) The second is highly ionized
   D) The second has more ionized hydrogen
22. The common ionic reaction of an acid with a base involves ions of
   A) Hydrogen and hydroxide ions
   B) Sodium and chloride
   C) Hydrogen and hydronium
   D) Hydroxide and nitrate

23. Which pH is a base solution?
   A) 3 B) 5 C) 7 D) 8
24. What is the concentration of nitric acid (HNO_3) solution with a pH of 2?
   A) 2M B) 0.1M C) 0.01M D) 0.2M
25. Base on Bronsted - Lowry concept, which of the following cannot functions as an acid?
   A) HSO_4^- B) H_2O C) NH_3 D) PCl_5
26. Which of the following is not alkalis?
   A) $Fe(OH)_3$ B) KOH C) $Ba(OH)_2$ D) NaOH
27. Which of the following represents an acidic solution ?
   A) pH = 1 C) $[H^+] = [A^-]$
   B) $[OH^+] = [HA]$ D) $[H^+] > [HA]$
28. Which of the following represents an acidic solution
   A) $[H^+] = 1.0 \times 10^{-7}M$ C) $[OH^-] = 2 \times 10^{-8}M$
   B) $[OH^-] = 5.0 \times 10^{-6}$ M D) $[H^+] = 4 \times 10^{-10}M$
29. The color of a solution of X turns to pink when a drop of phenolphthalien is added to it, solution X should be
   A) Acidic B) Basic C) neutral D) None
30. All of the following are basic. **EXCEPT**
   A) SnO B) Fe_2O_3 C) CuO D) Na_2O
31. A bottle contains a solution that is labeled pOH = 1. This solution is described as
   A) Weakly acidic C) Strongly acidic
   B) Weakly basic D) Strongl basic
32. Which one of the following statements is correct about the properties of acids?
   A) The pH of an aqueous solution of an acid increases as more water is added to it
   B) The pH of an acid increase with an increase in the concentration of an acid.

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C) Acids have bitter taste and change color of indicators.

D) Acids are electron pair and proton donors

33. Element X reacts with chlorine to form a compound with the formula XCl_2. The oxide of this element is basic. Element X is most likely

A) Ca B) Na C) P D) C

34. Which of the following oxides are acidic?

I. CaO III. NO_2 V. SO_2

II. Na_2O IV. CO VI. P_2O_5

A) III, V and VI C) I and II

B) III, IV, V and VI D) I, II and II

35. Which of the following substances is both a strong acid and strong oxidizing agent?

A) HCl B) H_2CO_3 C) H_3PO_4 D) H_2SO_4

36. Binary compounds that contain oxygen and any other element are:

A) Salts B) Acids C) Oxides D) Bases

37. Among the following, identify a substance that is capable of donating a pair of electrons

A) An Arrhenius acid C) An Arrhenius base

B) A Lewis acid D) A Lewis base

38. The base that is used as an electrolyte in some dry cells as to make soft soaps is

A) Calcium hydroxide C) Sodium hydroxide

B) Potassium hydroxide D) Ammonia

39. Which of the following reactions of chlorine gives an acidic solution?

A) $Cl_2(g) + 2Br^-(aq) \rightarrow$ C) $2Fe(s) + 3Cl_2(g) \rightarrow$

B) $Cl_2(g) + H_2O(l) \rightarrow$ D) $H_2(g) + Cl_2(g) \rightarrow$

40. What is the general name of compounds that contain cation derived from base and anion derived from acids?

A) Salts C) Acids

B) Oxides D) Bases

IV Short answer items

41. A student was given four unknown solutions. Each solution was checked for conductivity and tested with phenolphthalien. The results are shown in the table blow.

Solution	Conductivity	Color with phenolphthalien
1	Good	Colorless
2	Poor	Color less
3	Good	Pink
4	Poor	Pink

Based on the data with unknown solution could of a

a) NaOH? b) H_2SO_4? c) H_2O?

42. The table below shows typical pH values found in four parts of the human digestive system?

Part of human digestive system	pH
Mouth	6.7
Stomach	2.0
Pancreas	8.0
Small intestine	7.3

Which body parts have

a) acidic solution

b) basic solution

43. Given 500ml of a solution that contants 0.4mol NaOH, what is the molarity of hydrogen ions in this solution at 25°C?

44. How many milliliters 0.5M H_2SO_4 can be prepared from 24.5g of H_2SO_4?

45. Calculate the mass of 500cm^3 of 0.1M KOH solution

46. Complete the following reactions.

a) Acid + active metal →

b) Acid + Carbonate →

c) Sulphite + Acid →

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d) Base + Acid →

47. Write the components of oral rehydration salt (ORS)

48. Write one application of flame colors

49. Classify the following substances as macronutrient or micromstrient.

a) Potassium c) Nitrogen e) Carbon

b) Iron d) Cobalt f) Phosphorus

50. Write the two types of fertilizers.

Answer for Supplementary Exercise

Part I

1. True	3. False	5. False	7. True	9. False
2. False	4. True	6. True	8. True	10. False

Part II

11. E	13. A	15. B	17. F	19. C
12. D	14. H	16. J	18. I	20. G

Part III

21. B	26. A	31. D	36. C
22. A	27. D	32. A	37. D
23. D	28. C	33. A	38. B
24. C	29. B	34. A	39. B
25. D	30. A	35. D	40. A

Part IV

41. a. Solution 3 b. Solution 1 c. Solution 2

42. a. Mouth and stomach

b. Pancreas and small intestine

43. 0.8M

44. 500ml

45. 2.8g

46. a. Acid + active metal → salt + hydrogen

b. Acid + carbonate → salt + water + carbondioxide

c. Sulphite + acid → salt + sulphurdioxide + water

d. Base + acid → salt + water

47. Sodium chloride, sugar and water which is used as hydrating agent.

48. Used in fire works

49. a,c,e and f are macronutrients

b. and d are micronutrients.

50. Natural and synthetic fertilizers.

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UNIT 9

Energy Changes and Electro Chemistry

Introduction

- ☑ Chemical reactions occur by absorbing or releasing energy. These reactions can be redox or non - redox reactions.
- ☑ The energy changes in chemical reaction may be as change of enthalpy or from one form of energy in to another
- ☑ Energy changes occur in chemical reaction and electro chemistry.
- ☑ Chemical reactions occur by absorbing or releasing energy. These reactions can be redox or non - redox reactions.
- ☑ The energy changes in chemical reaction may be as change of enthalpy or from of energy into another.

9.1 Energy Changes in Chemical Reaction

- ☑ **Chemical reaction** is the process of converting reactants into products.
- ☑ The reaction vessel where the reaction takes place is known as system.

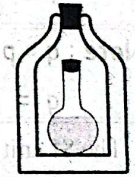
System is part of the universe selected for experimentation. It is divided into three

System

Open system
→ Exchanges matter and energy with its surrounding
Example
Decomposition of $CaCO_3$ in a crucible

Closed system
→ Exchanges energy but not matter with its surrounding
Example
A bottle of spirit kapt in refrigerator

Isolated system
→ Exchanges neither energy nor matter with its surrounding
Example
A reaction taking place in a thermos flask



Surrounding is every thing else out of the system

Universe is the sum of system and surrounding

System + surrounding = universe

Internal energy (E) is the total energy of system.

☑ It is the sum of potential energy and kinetic energy

$$E = P.E + K.E$$

☑ **Potential energy** is stored energy. It is a factor of :
- Attractive and repulsive interactions between molecules.
- Attractive interactions between electrons and nuclei molecules
- Repulsive interactions between electrons with in molecules
- Repulsive interactions between nuclei within molecules

☑ **Kinetic energy** is energy due to motion. It is component of internal energy consists of varous types of molecular motion and movement of electrons within molecules.

☑ When a system changes reactants to products and products to reactants,the internal energy has changed

$$\Delta E = E_f - E_i = E_{products} - E_{reactants}$$

☑ According to first law of thermodynamics, the total energy of the universe is conserved.

☑ The energy transferred between the system and the surrounding is the form of **heat** and **work** .

☑ **Heat (q)** is the transfer of energy between two bodies at different temperature.

Mathematically : $q = mC\Delta T$

☑ m - mass, ΔT = change in temperature $= T_f - T_i$

☑ ΔT is same in degree celcius and kelvin scale

☑ C = specific heat capacity

Note : q = positive implies heat gained by the system.

q = negative means heat lost by the system

☑ The SI unit of heat is joule (J) other unit is calorie

Example

$\boxed{1cal = 4.18\ J}$

How much heat is required to raise the temperature of 100g of water from 20 °C to 50 °C ?

Solution

$\Delta T = T_f - T_i = 50\,^{o}C - 20\,^{o}C = 30^{o}C = 30K$

$C = 4.18\ J/g.K, m = 100g$

$q = m\,C\,\Delta T = 100\ g \times 4.18\ J/g.K \times 30K = 12.5kJ$

Work (W) is energy transferred between a system and surrounding.

Note : W = positive means work done by the surrounding or work done On the system

W = Negative is work done by the system or work done on the Surrounding

Illustrative Examples

1. The total energy contained by the system is

A) Work
B) Heat
C) Heat energy
D) Internal energy

Answer : D

2. What type of energy change occurs when dynamite explodes ?

A) Chemical energy to kinetic energy
B) Kinetic energy to chemical energy
C) Chemical energy to electrical energy
D) Electrical energy to chemical energy

Answer : A

3. Which type of system exchanges both matter and energy with its surrounding ?

A) Open system
B) Closed system
C) Isolated system
D) None

Solution

A system that exchanges matter and energy with its surrounding is open system.

Answer : A

4. Which of the following is correct about the heat absorbed by 15 g of water to raise its temperature from 20 °C to 50 °C at constant pressure ?

A) 1.88 k J
B) 4.18 k J
C) 15 k J
D) 30 k J

Solution

$$q = m\,c\,\Delta T = 15g \times \frac{4.18J}{g.K} \times (50\text{-}20)K = 1881J = 1.88kJ$$

Answer : A

5. A hot dog provides 160 calories. What is this energy in joules ?

Solution

$$160cal = 160\ cal \times \frac{4.18J}{cal} = 668.8J$$

6. What is the change in internal energy of a system if it absorbed 300 kJ of heat from the surroundings and does 500 kJ of work ?

Solution

q = + 300 kJ because it is heat absorbed by the system

W = - 500 kJ because if is work done by the system.

$\Delta E = q + W = 300\ kJ - 500\ kJ = -200\ kJ$

7. The energy transferred between system and its surrounding occurs in two forms. What are these energies ?

Answer : Work and heat

8. System plus surrounding is ____

Answer : Universe

Exothermic and Endothermic Reaction

Based on absorption and release of energy, chemical reactions are classified into two. These are exothermic and endothermic reactions.

Table 9.1 The Difference Between Exothermic and Endothermic Reactions.

Exothermic reaction	Endothermic reaction
➥ Releases energy	➥ Absorbs energy
➥ Enthalpy change (ΔH)is negative ($\Delta H < 0$) $\Delta H = H_p - H_r$	➥ Enthalpy change is positive ($\Delta H > 0$)
➥ Heat content of product is less than that of reactant ($H_p < H_r$)	➥ Heat content of product is greater than that of reactant ($H_p > H_r$)
➥ Heat is added to the product side Reactant → product + heat	➥ Heat is added to the reactant side. Reactant + heat → product
➥ The reaction vessel becomes warm	➥ The reaction vessel becomes cold.
➥ Reactants are unstable	➥ Products are unstable
➥ Thermometer reading increases	➥ Thermometer reading decrease

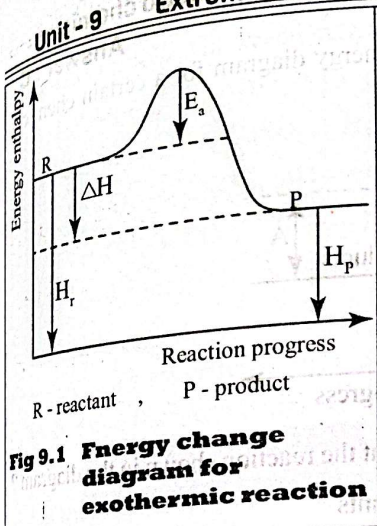


Fig 9.1 Energy change diagram for exothermic reaction

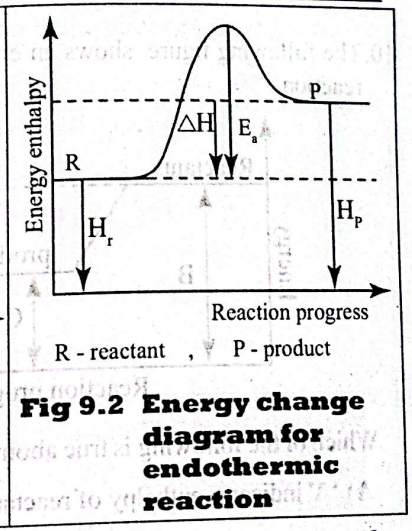


Fig 9.2 Energy change diagram for endothermic reaction

☑ **Activation energy (E_a)** is the minimum amount of energy required for a reaction to occur.

☑ **Calorimeter** is the device used to measure the amount of heat energy absorbed or released by a chemical reaction.

☑ **Heat capacity** is the quantity of heat transferred per changes in temperature.

Heat capacity = heat / change in temperature

☑ The heat capacity of one gram of a substance is called **its specific heat capacity (C)**

☑ The heat capacity of one mole of a substance is known as **molar heat capacity.**

Molar heat capacity = specific heat capacity × molar mass.

Illustrative Examples

9. Which of the following statements is false ?

A) The reaction vessel cools when an endothermic reaction occurs.

B) An exothermic reaction is characterized by a negative value of ΔH.

C) An endothermic reaction causes the surrounding to absorb heat

D) Heat is evolved when an exothermic reaction occurs.

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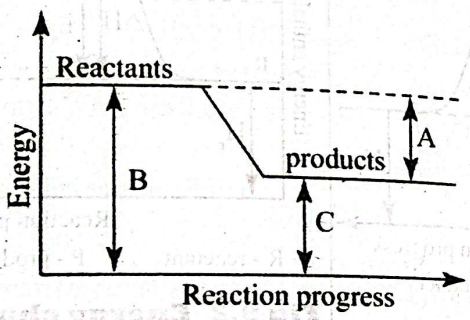
10. The following figure shows an energy diagram for a certain chemical reaction.



Which of the following is true about the reaction shown in the diagram?

A) 'A' indicates enthalpy of reactants

B) 'B' indicates heat change of reaction

C) 'C' indicates heat content of product

D) The reaction is endothermic

Solution

➥ $A = \Delta H$, $B = H_r$ and $C = H_p$

➥ The reaction is exothermic because $H_r > H_p$

Answer : C

11. Assume that the energy of the products in a chemical reaction is 50 kcal. This reaction could be exothermic if the energy of the reactants was :

A) 50 kcal B) 20 kcal C) 60 kcal D) 10 kcal

Solution

For exothermic reaction, $\Delta H < 0$

$\Rightarrow \Delta H = H_p - H_r < 0$

$\Rightarrow H_p < H_r \Rightarrow H_r > H_p$. Thus, the answer must be above 50 kcal.

Answer : C

12. Which types of chemical reaction generates heat energy ?

Answer : Exothermic reaction

13. Classify the following reactions as endothermic or exothermic reactions.

a) $CaCO_3 \xrightarrow{heat} CaO + CO_2$

b) $Ca + H_2SO_4 \rightarrow CaSO_4 + H_2 + heat$

c) $N_2 + O_2 \rightarrow 2NO_2$, $\Delta H = 66.4kJ$

d) $C + O_2 \rightarrow CO_2$, $\Delta H = -394$ kJ

e) Photosynthesis f. Respiration

Solution

➥ a, c and e are endothermic reactions.

➥ b, d and f are exothermic reactions.

14. Write the two importance of chemical changes or chemical reactions.

Solution

1. Production of new substances
2. Production of heat energy

15. Write the difference between

a) Endothermic reaction and endothermic process

b) Exothermic reaction and exothermic process.

Solution

a) **Endothermic reaction** is the chemical change that absorbs heat but an endothermic process is a chemical or physical change that absorbs heat.

➥ Melting, evaporation and sublimation are physical changes that absorb heat.

b) **Exothermic reaction** is chemical change that releases energy however, exothermic process is either a physical or chemical change that releases heat.

➥ Freezing, condensation and deposition are physical changes that release heat.

16. How could you identify whether a reaction is endothermic or exothermic in the laboratory ?

Solution

By using our skin. If we feel hot it is exothermic but if we feel cold,

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the reaction is endothermic.

17. Calculate the molar heat capacity of water.

Solution

$$\text{Molar heat capacity} = C \times M_{water},\ C = 4.18 \frac{J}{g.K}$$

$$= 4.18 \frac{J}{g.K} \times 18 \frac{g}{mol} = 75.24 \frac{J}{mol.K}$$

9.2 Energy Changes and Electro Chemistry

☑ **Electro chemistry** constitutes one of the most important interfaces between chemistry and every day life. Every time you start your car, turn on your calculator, look at your digital watch or listen to radio at the beach, you are depending on electro chemical reactions. Batteries are the result of electro chemistry.

☑ **Electro chemistry** is defined as the study of the interchange of chemical and electrical energy.

☑ It is primarily concerned with two processes that involve oxidation reduction reactions : the generation of an electric current from a chemical reaction and the use of a current to produce chemical change.

☑ The apparatus where this energy conversion takes place is called electro **chemical cell.**

Electro chemical cell

Electrolytic cell	Galvanic cell
→ Converts electrical energy to chemical energy	→ Converts chemical energy to electrical energy

Electrical Conductivity

☑ It is the ability of a substance to transmit electricity.

☑ The substances that transmit electricity are called **electrical conductors.**

☑ There are two types of conduction. These are metallic and electrolytic **conductions.**

Table 9.2 The Difference Between Metallic and Electrolytic Conditions.

Metallic (electronic) conduction	Electrolytic / ionic / condition
➥ Are also called electronic conductors because the charge carriers are mobile electrons ➥ Transmit electricity without undergoing chemical change . ➥ Conduct electricity in all states **Example** Metals	➥ Are also known as ionic conductors because the charge carriers are mobile negative and positive ions. ➥ Transmit electricity by undergoing chemical change. ➥ Conduct electricity in aqueous or molten states **Example** Salts, acids and bases

Strong electrolytes - ionize almost completely in water

➥ Produce large amount of ions that carry charge.

Example

Strong acids, strong bases, molten and soluble salts.

Weak electrolytes : ionize very slightly in water

➥ Produce small amounts of ions when dissolved in water.

Example

Weak acids, weak bases, partially soluble salts.

Non - electrolytes : are solutions that do not transmit electricity due to the absence of charge carriers.

Example

Sugar solution, ethanol solution, non - polar compounds in non - polar solvents.

Conducting Apparatus

It is the instrument used to test the conductivity of substances.

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Its components are :

- Connecting wire
- Electrodes
- Light bulb
- Switch
- Beaker

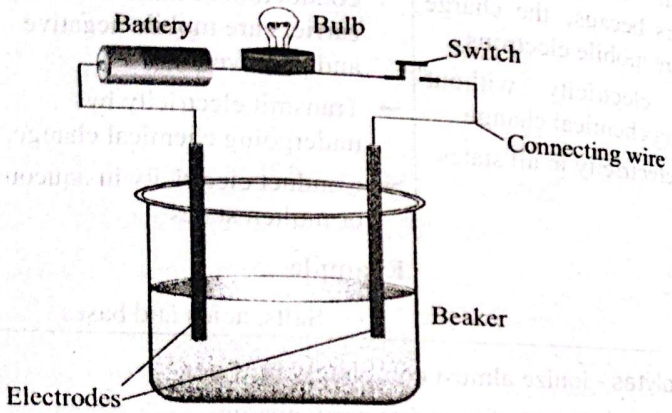

Fig 9.3 Conducting Apparatus

Illustrative Examples

18. Which of the following is weak electrolytes ?

A) Sulphuric acid
C) Nitric acid
B) Acetic acid
D) Hydrochloric acid

Solution

Strong acids are strong electrolytes and weak acids are weak electrolyes.

Answer : B

19. The capacity of a substance to transmit electricity is termed as

A) Conductivity
C) Resistivity
B) Fluidity
D) Permittivity

Answer: A

20. What are weak electrolytes ?

A) They are ionic and polar covalent compounds.

B) They release limited number of ions in water

C) They are strong acids and bases.

D) They include sugar, alcohols and hydrocarbons

Answer : B

21. Of the following substances one is an electrical conductor in solid state. Which one is it ?

A) Sodium chloride
C) Potassium hydroxide
B) Sulphur crystal
D) Aluminum

Solution

Metals transmit electricity but not salts and bases because the ions are not free to move.

Answer: D

22. Which of the following are major charge carriers in electrolytic conduction ?

A) Water molecules
C) Cations and anions
B) Freely moving electrons
D) Neutral atoms.

Solution

The charge carriers in electrolytic conditions are positive and negative ions.

Answer : C

23. Which of the following conduct electricity without undergoing any chemical change ?

A) Sulphuric acid solution
C) Copper metal
B) Sodium chloride solution
D) Molten lead bromide

Solution

Metallic conductors transmit electricity without undergoing chemical change.

Answer : C

24. Which of the following does not transmit electricity ?

A) KCl (l)
B) $COCl_2$ (aq)

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C) Aluminum D) $MgF_2(s)$

Answer : D

25. Classify the following as strong electrolyte, weak electrolyte or non-electrolyte

a) Ammonia solution
b) Ethanol solution
c) Sodium nitrate solution
d) Hydrofluoric acid solution
E) Molten potassium chloride
F) Solid calcium chloride

Solution

- a and d are weak electrolytes
- b and f are non - electrolytes
- c and e are strong electrolytes

26. The apparatus used to test conductivity of substances is known as _.

Answer : Conducting apparus

27. What are the two types of electrochemical cells ?

Answer : Electrolytic and galvanic cells

9.3 Electrolytic Cell

☑ It is an electro chemical cell that converts electrical energy to chemical energy.

☑ Electricity causes non - spontaneous reaction to occur.

Its major components are :

- Connecting wire
- Electrodes
- Direct current source
- Electrolyte

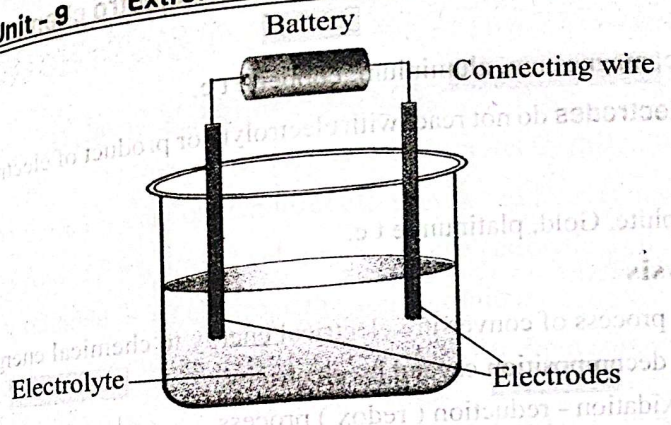

Fig 9.4 Electrolytic cell

Electrodes

- Are conducting rods.
- Conduct electrons into and out of an electrolyte
- Are often a metal plate or wire or a piece of graphite

Classification of Electrodes

Based on the half - reaction, electrodes are classified as

1. **Anode** is electrode where oxidation half reaction occurs

Use **Anox** to rember it easily. (An - Anode, ox - oxidation)

- It electrode where electrons leave the solution.
- It is positive in electrolytic cell and negative in galvanic cell.

2. **Cathode** is electrode where reduction half reaction occurs.

Use **car** to rember it easily (Ca - cathode r - reduction)

- It is electrode where electron enter the solution.
- It is negative in electrolytic cell and positive in galvanic cell.

Based on their Nature, Electrodes are Classified as

1. **Active electrodes** react with electrolyte or product of electrolysis or themselves oxidized.

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Example

Zinc, magnesium, aluminium, copper e t c.

2. Inert electrodes do not react with electrolyte or product of electrolysis.

Example

Graphite, Gold, platinum e t c.

Electrolysis

- It is the process of converting electrical energy to chemical energy.
- It is the decomposition caused by electricity.
- It is a oxidation - reduction (redox) process
- It takes place in electrolytic cell.

Electrolysis of Molten or Fused Electrolytes

- The electrolyte decomposes, so electrical energy causes a chemical change
- **Cations** migrate toward cathode and discharged by gaining electrons.
- **Anions** move toward anode and discharged by losing electrons.
- Electrons flow from anode to cathode in the external circuit.
- Battery is a driving force for electrons to move from anode to cathode.
- Electrons are released at anode and consumed at cathode
- Electrons lost at anode are equal to electrons gained at cathode, in cell or overall reaction.

Example

Electrolysis of molten or fused lead (II) bromide,

$Pb^{2+} + 2Br^- \xrightarrow{\text{melting}} Pb^{2+} + 2Br^-$

$\Rightarrow Pb^{2+}$ moves toward cathode and discharged there.

$\Rightarrow Br^-$ migrates toward anode and discharged there

Cathode reaction : $Pb^{2+} + 2e^- \rightarrow Pb$ ------ reduction

Anode reaction : $2Br^- \rightarrow Br_2 + 2e$ ---- oxidation

Cell reaction : $Pb^{2+} + 2Br^- \xrightarrow{\text{Electrolsis}} Pb + Br_2$ --- Redox reaction

Illustrative Example

28. Which of the following combinations is correctly matches ?

A) Cation → an ion that loses electron / s / to form atoms

B) Anode → a positive electrode in an electrolytic cell

C) Cathode → an electrode that attracts anions.

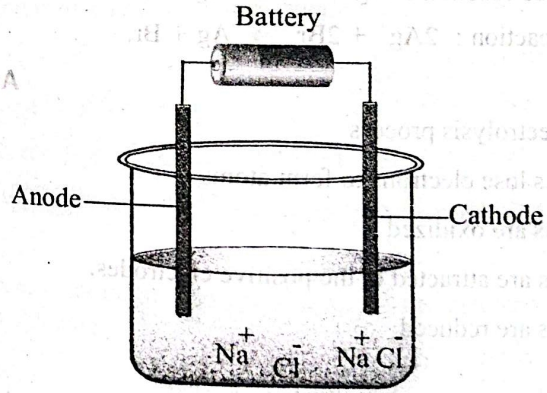
D) Anion → an ion that gains electron / s / to from atoms.

Solution

Cations gain and anions lose electrons to form atoms at cathode and anode respectively.

Answer : B

29. The following diagram shows the electron / s / of molten sodium chloride, using an electrolytic cell.



Which statements is true about this cell ?

A) Electrons flow from cathode to anode

B) Reduction takes place at the positive electrode

C) Chloride is discharged at the anode

D) Sodium ion is discharged at the anode

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Solution

- Electrons flow from anode to cathode
- Reduction occurs at negative electrode
- Chloride is discharge at anode and sodium ion is discharged at cathode

Answer : C

30. Which of the following represents the reduction half - reaction for the electrolysis of molten silver bromide ?

A) $2Br^- \rightarrow Br_2 + 2e$ C) $Ag^+ + e^- \rightarrow Ag$

B) $Br_2 + 2e^- \rightarrow 2Br^-$ D) $Ag \rightarrow Ag^+ + e^-$

Solution

$AgBr \xrightarrow{\text{melting}} Ag^+ + Br^-$

Exnode reaction : $2Br^- \rightarrow Br_2 + 2e^-$

Cathode reaction : $Ag^+ + e^- \rightarrow Ag$

Cell reaction : $2Ag^+ + 2Br^- \rightarrow Ag + Br_2$

Answer : C

31. During electrolysis process

A) Cations lose electrons to form atoms

B) Cations are oxidized

C) Anions are attracted to the positive electrodes.

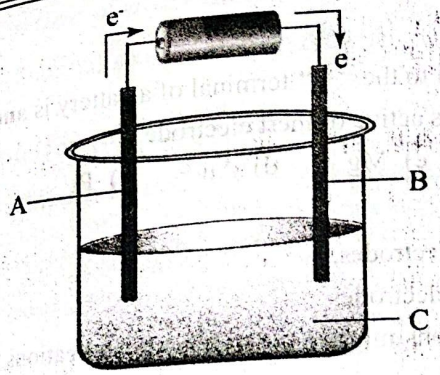
D) Anions are reduced

Solution

Cations are reduced and anions are oxidized during electrolysis.

Answer : C

32. The diagram below shows electrolysis of molten $PbCl_2$ using chemically inert electrode



Which of the following is correct ?

A) Electrons flow from cathode to anode

B) Lead ion acts as a reducing agent

C) Chloride ion is reducing agent

D) Lead ion is discharged at the cathode.

Solution

$PbCl_2 \rightarrow Pb^{2+} + 2Cl^-$

Anode reaction : $2Cl^- \rightarrow Cl_2 + 2e^-$

Cathode reaction : $Pb^{2+} + 2e^- \rightarrow Pb$

Answer : D

33. Write anode, cathode and cell reactions for the electrolysis of molten potassium chloride.

Solution

$KCl \xrightarrow{\text{melting}} K^+ + Cl^-$

Anode reaction : $2Cl^- \rightarrow Cl_2 + 2e^-$

Cathode reaction : $K^+ + e \rightarrow K$

Cell reaction : $2K^+ + 2Cl^- \rightarrow 2K + Cl_2$

34. Two steel needles dipped into dilute hydrochloric acid solution are connected to the terminal of a battery. What is the needle connected to

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the end of the battery marked with a "+" ?

Solution

The electrode connected to the "+ " terminal of a battery is anode.

35. Classify the following as active or inert electrode

a) Cu b) Ag c) Mg d) Au e) Pt f) Zn

Solution

- b, d and e are inert electrodes.
- a, c and f are active electrodes.

36. During electrolysis, anions migrate toward anode and cations migrate toward anode. Why ?

Solution

This is because of unlike charge attract each other.

37. What is the difference between half - reactions and cell reaction ?

Solution

Half - reactions occur at electrodes, but cell reaction is the combination of half - reactions.

38. The electrodes that do not themselves involve in the redox reactions are called ____

Answer : Inert electrodes

39. Why electrolysis is a redox process ?

Solution

Because it is the combination of oxidation half reaction and reduction half - reaction.

Application of Electrolysis

Industrially, electrolysis has the following applications.

1. Production of Active Metals

- Group I A and I I A metals are active metals
- Aluminum is also manufactured industrially by Hall process using electrolysis

2. Production of non - metal

Chlorine, fluorine, oxygen and hydrogen are industrially produced by electrolysis.

3. Production of Compounds

Sodium hydroxide and potassium hydroxide are commercially produced from electrolysis of concentrated sodium chloride and potassium chloride solution respectively.

4. Electroplating

It is the process that involves the coating of a solid surface with metal layer by the application of electrolysis.

During electroplating :

- Anode = plating metal
- Cathode = material to be plated
- Electrolyte = soluble or molten salt of plating metal.

Example

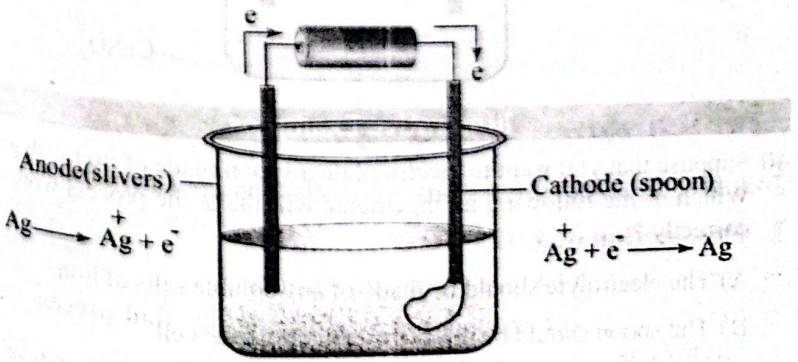
Electroplating of spoon made of ion with silver.

Solution

Anode = silver

Cathode = spoon

Electrolyte = $AgNO_3$



5. Electrorefing of Metals

It is purification of metals using electrolysis.

During electro refing:

Anode = impure metal

Cathode = pure metal

Electrolyte = salt of the metal.

Example

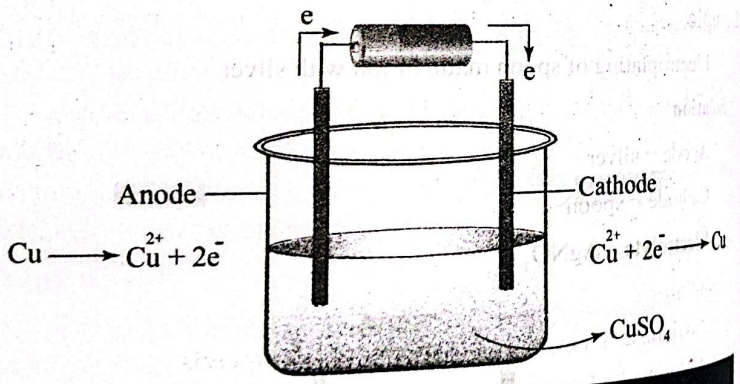
Electrorefing of copper :

Anode = Impure Cu

Cathode = pure Cu

Electrolyte = $CuSO_4$

Note : Mass of anode decreases but mass of cathode increases



Illustrative Examples

40. Suppose that you want to electro plate a spoon made of iron with gold. Which of the following arrangements will make the process work out correctly ?

A) The electrolyte should be made of any soluble salts of iron.

B) The spoon should be made the cathode of the cell

C) The electrolyte can be prepared from soluble salts of metals other than gold.

D) The cell should be constructed as a simple galvanic cell

Solution

➥ Anode = gold, cathode = spoon, electrolyte = soluble salt of gold.

➥ Electrochemical cell = electrolytic cell

Answer : B

41. Which of following is not the industrial applications of electrolysis ?

A) Production of metals

B) Production of non - metals

C) Electroplating

D) Production of electricity

Answer : D

42. Which of the following is true about electrolytic refining of copper ?

A) The setting impurity may contain gold, silver and platinum

B) The nature of electrolyte changes significantly

C) The pure copper deposits at the anode

D) The impure copper is used as cathode.

Solution

Anode = Impure Cu, Cathode = pure Cu, Electrolyte = $CuSO_4$

Answer : A

43. List some metals, non - metals and compounds that can be manufured by electrolysis.

Solution

➥ Sodium, calcium, magnesium, aluminum: **metals**

➥ Chlorine, hydrogen, oxygen, fluorine: **non - metals**

➥ Sodium hydroxide, potassium hydroxide: **compounds.**

44. What is the relationship between amount of metal plated and amount of electricity delivered during electro plating ?

Solution

The amount of metal plated is directly related to the amount of electricity delivered

45. During electrorefing, why the mass of cathode increases ?

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Solution

Reason : Pure metal is deposited on it .

46. The process of coating one metal with a thin layer of another metal using electricity is called ____

Answer : Electroplating

9.4 Galvanic / voltaic / Cells

Galvanic cells are electrochemical cells that convert chemical energy to electrical energy

- In galvanic cells, spontaneous chemical reactions generate electricity.
- Anode is negative and cathode is positive

There are three types of galvanic cells. These are

1. Primary galvanic cells
2. Secondary galvanic cells
3. Fuel cells

I. Primary Galvanic Cells

They are not recharged because the cell reaction is irreversible

Some of the primary cells are :

i. The zinc - copper cell (Daniell cell)

- ☑ It is wet galvanic cell because the electrolyte is aqueous solution
- ☑ Anode = Zinc in $ZnSO_4$ Solution
- ☑ Cathode = Copper in $CuSO_4$ Solution
- ☑ The anode and cathode compartments are separated by an inverted U - tube filled with neutral salt solution called **salt bridge.**
- ☑ Anode reaction : $Zn \rightarrow Zn^{2+} + 2e^-$
- ☑ Cathode reaction : $Cu^{2+} + 2e \rightarrow Cu$
- ☑ Cell reaction : $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$, voltage (V) = 1.10 V
- ☑ As the zinc ions leave the solid zinc anode and enter the solution, electrons are lost behind the zinc anode, as a result, the zinc anode electrode carries a **negative charge.**
- ☑ At the copper cathode, Cu^{2+} ions become attached to the electrode and

seeks electrons to become reduced. This gives the copper electrode a **positive charge**

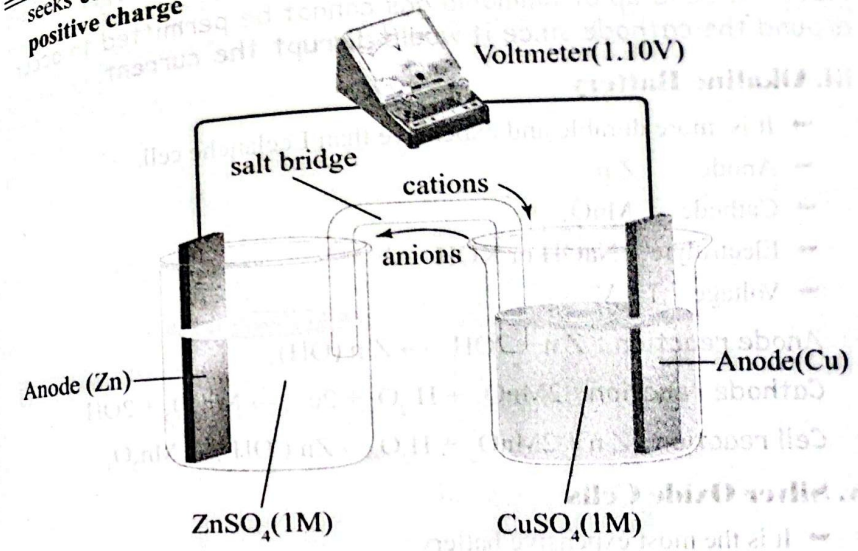

Fig 9.5 Daniel Cell

A salt Bridge Serves Three Functions

1. It allows electrical contact between the two solutions.
2. It prevents mixing of the electrode solutions
3. It maintains the electrical neutrality in each half - cell as ions flow into and out of the salt bridge.

ii. The Leclanche Dry Cell (zinc - carbon cell)

Anode = zinc

Cathode = graphite

Electrolytes = NH_4Cl, $ZnCl_2$, MnO_2

Voltage = 1.5 V

Anode reaction : $Zn \rightarrow Zn^{2+} + 2e$

Cathode reaction : $2NH_4^+ + 2MnO_2 + 2e \rightarrow Mn_2O_3 + 2NH_3 + H_2O$

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Cell reaction : $Zn + 2NH^{+}_{4} + 2MnO_2 \rightarrow Zn^{2+} + Mn_2O_3 + H_2O$

Note : A build up of ammonia gas cannot be permitted to occur around the cathode since it would disrupt the current.

iii. Alkaline Battery

- It is more durable and expensive than Leclanche cell.
- Anode = Zn
- Cathode = MnO_2
- Electrolyte = NaOH or KOH
- Voltage = 1.5 V.

Anode reaction : $Zn + 2OH^- \rightarrow Zn(OH)_2$

Cathode reaction : $2MnO_2 + H_2O + 2e^- \rightarrow Mn_2O_3 + 2OH^-$

Cell reaction : $Zn + 2MnO_2 + H_2O \rightarrow Zn(OH)_2 + Mn_2O_3$

iv. Silver Oxide Cells

- It is the most expensive battery
- Anode = Zn
- Cathode = Ag_2O
- Electrolyte = NaOH or KOH
- Voltage = 1.5 V.

Anode reaction : $Zn + 2OH^- \rightarrow Zn(OH)_2 + 2e^-$

Cathode reaction : $Ag_2O + H_2O + 2e^- \rightarrow 2Ag + 2OH^-$

Cell reaction : $Zn + Ag_2O + H_2O \rightarrow Zn(OH)_2 + 2Ag$

v. Mercury cells

- Anode = Zn
- Cathode = HgO
- Electrolyte = NaOH or KOH
- Voltage = 1.35 V

Anode reaction : $Zn + 2OH^- \rightarrow Zn(OH)_2 + 2e^-$

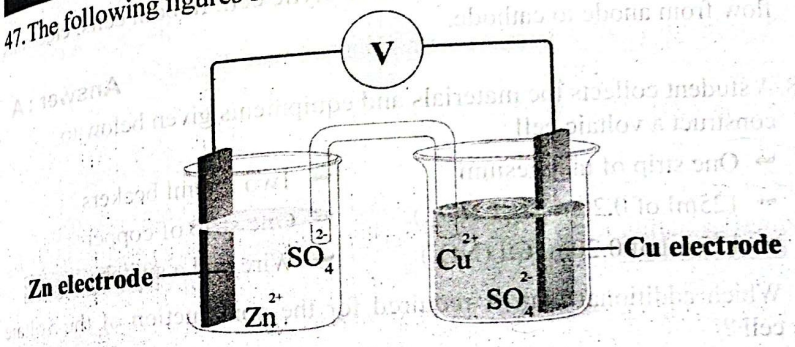
Cathode reaction : $HgO + H_2O + 2e^- \rightarrow Hg + 2OH^-$

Cell reaction : $Zn + HgO + H_2O \rightarrow Zn(OH)_2 + Hg$

Note : Mercury cells pose greater harm to the environment because mercury and mercury compounds are poisonous.

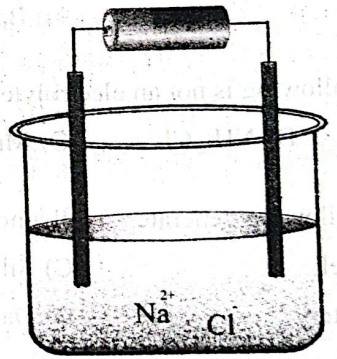
Illustrative Examples

47. The following figures show electrochemical cells.



Cell A

battery



Cell B

Which of the following statements is true about these cells ?

A) Cell A converts chemical energy into electrical energy while cell B converts electrical energy into chemical energy.

B) In cell B electrons flow from the negative terminal to the positive terminal of the battery in the external circuit.

C) In cell A electrons flow from Cu - electrode to Zn - electrode

D) Cell A Converts electrical energy into chemical energy while cell B converts chemical energy into electrical energy.

Solution

Cell A is galvanic and cell B is electrolytic cell. In both cells, electrons flow from anode to cathode.

Answer : A

48. A student collects the materials and equipments given below to construct a voltaic cell

- One strip of magnesium
- 125ml of 0.20m Mg (NO_3)$_2$
- 125 ml of 0.20m Cu (NO_3)$_2$
- Two 250ml beakers
- One strip of copper
- Wire and a switch

Which additional item is required for the construction of the voltaic cell ?

A) Cathode C) Anode

B) Salt bridge D) Battery

Answer : B

49. Which of the following is not an electrolyte in leclanche cell ?

A) H_2SO_4 B) NH_4Cl C) MnO_2 D) $ZnCl_2$

Answer : A

50. Which of the following generates small amount of voltage ?

A) Leclanche cell C) Silver oxide cell

B) Alkaline battery D) Daniel cell

Solution

Most primary galvanic cells generate 1.5 V but Daniel cell generates 1.10 V.

Answer : D

51. The anode used in most primary galvanic cells is ____.

Answer : Zinc

52. Classify the following primary galvanic cells as dry cell or wet cell.

a) Daniel cell c) Alkaline battery

b) Leclanche cell d) Silver oxide cell

Solution

- Galvanic cells with aqueous electrolytes are wet cells while those with paste electrolytes are dry cells.
- b, c and d are dry cells.
- a is wet cell.

53. Identify the negative and positive electrodes in leclanche cell.

Solution

Negative electrode = zinc

Positive electrode = graphite

54. How a salt bridge maintains electrical neutrality between the two electrodes solutions ?

Solution

The salt bridge releases negative ion into anode compartment and positive ion into the cathode compartment to maintain electrical neutrality between the two solutions.

55. What problem initiated scientists to discover dry cells ?

Solution

Because galvanic cells with aqueous electrolytes are not easily portable from place to place.

2. Secondary Galvanic cells / Reversible cells /

They are rechargeable galvanic cells because the cell reactions are reversible.

i. Nickel - cadmium battery (Nicad battery)

- It is used in battery - operated tools and portable computer.
- Anode = cadmium (Cd)
- Cathode = Nickel (IV) oxide (NiO_2)
- Electrolyte = NaOH or KOH

Anode reaction : $Cd + 2OH \rightleftharpoons Cd(OH)_2 + 2e$

Cathode reaction : $NiO_2 + 2H_2O + 2e^- \rightleftharpoons Ni(OH)_2 + 2OH^-$

Cell reaction : $Cd + NiO_2 + 2H_2O \rightleftharpoons Cd(OH)_2 + Ni(OH)_2$

- Nicad battery has a voltage of 1.40 V.
- It has longer life than a lead storage battery
- It can be packed and sealed easily.

ii. Lead Storage Battery /car battery /

- It is used in most automobiles.
- Anode = Pb
- Cathode = PbO_2
- Electrolyte = H_2SO_4

Anode reaction : $Pb(s) + SO_4^{2-}(aq) \rightleftharpoons PbSO_4(s) + 2e^-$

Cathode reaction : $PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^- \rightleftharpoons PbSO_4(s) + 2H_2O(l)$

Cell reaction : $Pb(s) + pbO_2(s) + 4H^+(aq) + 2SO_4^{2-}(aq) \rightleftharpoons 2PbSO_4(s) + 2H_2O(l)$

3. Fuel cells

- They are galvanic cells that require continuous supply of reactants to keep functioning.
- Fuel cells do not store chemical energy.
- The common fuel cells is hydrogen - oxygen fuel cell.

Illustrative Examples

56. Which of the following is an example of a secondary voltaic cell?

A) Zn - C dry cell C) Daniel cell

B) Lead storage battery D) Leclanche cell

Answer : B

57. Which of the following electro chemical cell is used in auto mobiles?

A) Nicad battery C) Daniel cell

B) Fuel cell D) Lead - storage battery

Answer : D

58. Consider the following electrochemical reaction in lead storage battery:

$Pb + PbO_2 + 2H_2SO_4 \rightarrow 2PbSO_4 + 2H_2O$

Which species is reduced when the battery is discharging ?

A) $PbSO_4$ B) PbO_2 C) H_2SO_4 D) Pb

Solution

The oxidation number of Pb in PbO_2 is + 4 while in $PbSO_4$ is + 2.
Thus, the reduced substance is PbO_2.

Answer : B

59. Write the anode, cathode and cell reaction during charging of a car battery

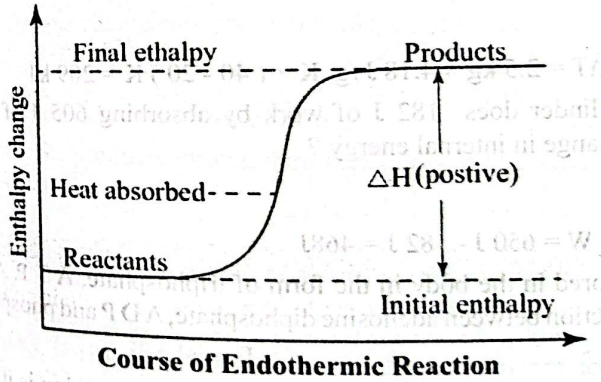
Solution

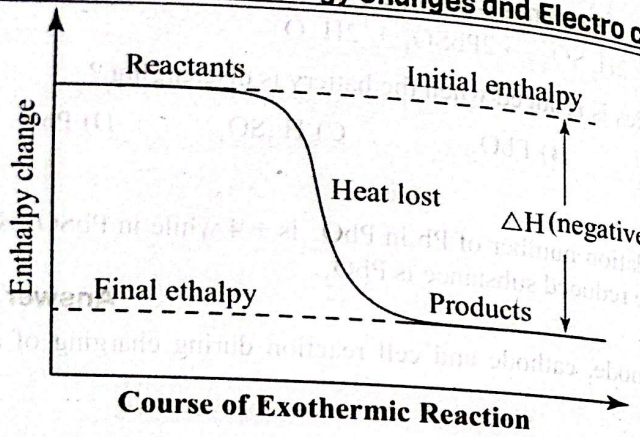
Anode reaction : $PbSO_4 + 2H_2O \rightarrow PbO_2 + 4H^- + SO_4^{2-} + 2e^-$

Cathode reaction : $PbSO_4 + 2e^- \rightarrow Pb + SO^{2-}{}_4$

60. Draw energy diagram for exothermic and endothermic reactions.

Solution





61. Write three changes of states that are exothermic.

Solution

Freezing, condensation and deposition

62. How much heat, in kilojoules, is required to raise the temperature of 2.5kg of water from 20°C to 40°C ?

Solution

$q = m\ c\ \Delta T = 2.5\ kg \times 4.18\ J/g\ K \times (40 - 20)\ K = 209\ kJ$

63. A gas in cylinder does 182 J of work by absorbing 605 J of heat. Calculate change in internal energy ?

Solution

$\Delta E = q + W = 650\ J - 182\ J = 468J$

64. Energy is stored in the body in the form of triphosphate, A T P. A T P forms on reaction between adenosine diphosphate, A D P and phospheric acid .

$A\ D\ P + H_3PO_4 + 38\ k\ J \rightarrow A\ T\ P + H_2O$. What type reaction is it ?

Answer : Endothermic reaction because heat is added to the reactant side.

65. Write the difference between electrolytic and galvanic cells.

Solution

Electrolytic cell	Galvanic cell
➥ Converts electrical energy to chemical energy	➥ Converts chemical energy to electrical energy
➥ Non - spontaneous redox reaction occurs.	➥ Spontaneous redox reaction generates electricity
➥ Anode is positive	➥ Anode is negative
➥ Cathode is negative	➥ Cathode is positive
➥ Consumes energy	➥ Generates energy.

66. For electrolysis of fused $CuCl_2$.

a) Identify the charge carriers
b) Write anode half - reaction
c) Write cathode half - reaction
d) Write the cell reaction

Solution

a) Cu^{2+} and Cl^-
b) $2Cl^- \rightarrow Cl_2 + 2e$
c) $Cu^{2+}\ 2e \rightarrow Cu$
d) $Cu^{2+} + 2Cl^- \rightarrow Cu + Cl_2$

67. What are the charge carriers in metallic conductors ?

Solution

Free electrons or valence electrons

68. What are the products formed during electrolysis of molten magnesium bromide ?

Solution

Mg and Br_2

69. A voltage cell is constructed with Pb (s) in Pb (NO_3)$_2$ (aq) Zn in Zn (NO_3)$_2$ (aq) for other cell. If Zn is a better reducing agent than Pb and the salt bridge contains an aqueous solution of KNO_3, answer the following questions.

a) Identify the anode and cathode
b) Which ions move towards the anode
c) Which ions move towards the cathode
d) Which metal is positive electrode
e) Draw the voltaic cell

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Solution

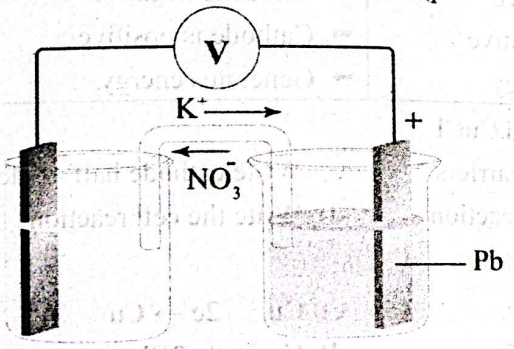
a) More active metal is anode and less active is cathode

Anode = Zn, cathode = Pb

b) NO^-_3 from the salt bridge moves into anode to neutralize the excess Zn^{2+} formed.

c) K^+ moves into cathode compartment to replace the charge of lead ion (Pb^{2+}) reduced. d. load (pb

d)



$Zn(NO_3)_2(aq)$ $Pb(NO_3)_2(aq)$

70. For electrolysis of fused sodium bromide, write the product formed at anode and cathode.

Solution

Anode reaction : $2Br^- \rightarrow Br_2 + 2e^-$

Cathode reaction : $Na^+ + e^- \rightarrow Na$

$\Rightarrow$ Na is formed at cathode and Br_2 is given off at anode.

71. Explain the difference between metallic and electrolytic conduction.

Solution

Metallic conduction

- No chemical change
- Transmit electricity in all states
- Free electrons are charge carriers

Electrolytic conduction

- There is chemical change.
- Transmit electricity in molten or aqueous state.
- Mobile ions are charge carriers.

Supplementary Exercise

I. True/ False Items

1. All chemical reactions accompany energy change
2. Internal energy is the total energy contained by a system.
3. Enthalpy change is calculated as enthalpy of reactant minus us that of product.
4. Electro chemistry is the study of energy change in redox reactions.
5. Non - electrolytes are electrical conductors that transmit electricity.
6. Electrolysis is the decomposition of an electrolyte using electrical energy.
7. Primary galvanic cells are rechargeable cells that are reused.
8. Anode is electrode attached to the negative terminal of the battery.
9. In voltaic cells, the polarity or charge of the electrodes is determined by the redox reactions.
10. In electrolytic cells the polarity of the electrodes is determined by the external power source.

II. Multiple choice Items

11. A ΔH of - 100 k J / mol indicates that the reaction is

A) Endothermic C) In need of catalyst

B) Unstable D) Exothermic

12. Enthalpy is an expression for the

A) Heat content C) Reaction rate

B) Energy state D) Activation energy

13. 100 cal of heat is equal to _____ joule of heat.

A) 100 J B) 18 J C) 418 J D) 41.8 J

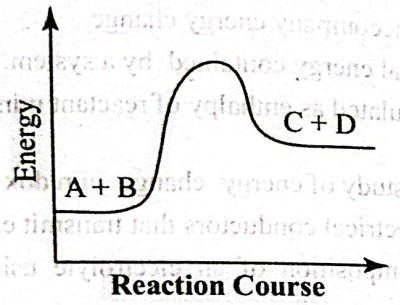
14. The enthalpies associated with the reaction, A + B → C + D is shown below. The fact that the enthalpy of C + D is higher than that of A + B tell us that

A) The reaction is exothermic

B) The reaction is endothermic

C) Heat content of reactant is higher than product

D) Heat content of product is lower than reactant.



15. A system that cannot exchange matter and energy with its surrounding is

A) Open system C) Isolated system

B) Closed system D) Homogeneous system

16. Which of the following when dissolved in water and placed in the conductivity apparatus would cause the light to glow ?

A) Table salt C) Sugar

B) Ethyl alcohol D) Glycerine

17. In the electrolysis of molten copper (II) chloride, the substance librated at the anode is

A) Copper C) Hydrogen

B) Chlorine D) Copper chloride

18. Salts will conduct an electric current when they are

A) Solidified C) Frozen

B) Melted D) Dehydrated

19. The cathode in an electro chemical cell is

A) Always negative C) Reduction takes place

B) Always positive' D) Oxidation takes place

20. Which of the following is true about electrolysis ?

A) It converts chemical energy to electrical energy.

B) Electrodes block flow of electrons during electrolysis.

C) Cathode attracts cations during electrolysis

D) Electrolyte is non - redox process

21. Which of the following is wet cell ?

A) Leclanche cell C) Nicad battery

B) Car battery D) Alkaline battery

22. The electrolyte in car battery is

A) H_2SO_4 B) HNO_3 C) HCl D) NaOH

23. All of the following are strong electrolytes. **Except**

A) HNO_3 B) KOH C) NH_3 D) $HClO_4$

24. The difference between a galvanic and electrolytic cell is that unlike electrolytic cell the galvanic cell :

A) Is used in electroplating process

B) Is used in production of metals from their ores

C) Uses electrical energy to bring about chemical changes

D) Has spontaneous redox reactions

25. An electrolytic cell contains molten copper (11) chloride. When direct current passes through the electrolyte the reaction below takes place.

$CuCl_2(1) \rightarrow Cu + Cl_2$

Which of the following is correct about this reaction ?

A) Electrons enter into the electrolyte at the anode

B) Electrons are takes up by the negative electrode

C) Chlorine gas is librated at the anode

D) Reduction occurs at positive electrode

26. In galvanic cell, oxidation occurs at

A) Anode B) Cathode

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C) Salt bridge D) Electrolyte

27. When a lead storage battery charged, lead (11) sulphate is
   A) Formed at cathode
   B) Formed at anode
   C) Formed at both electrodes
   D) Removed from both electrodes

28. Increasing the concentration of ions in an electrolyte Solution :
   A) Increase the extent of conduction through it
   B) Decreases the extent of conduction of through it
   C) Has no effect on conduction of electricity
   D) Changes the direction of electron flow

29. You want to plate a medal made of copper with gold, your anode should be
   A) Gold B) Medal C) Copper D) None

30. Metals are produced by
   A) Oxidation C) Redox reaction
   B) Reduction D) Non - redox reaction

III. Short answer items

31. ____ is a solution or a molten compound which cannot be decomposed by an electric current.

32. ___ are two poles of graphite or metal at which electrons enter or leave an electrolyte.

33. Given the cell reaction : $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$. Write the anode and cathode half - reactions.

34. A system absorbs 1000 cal of heat and also does 2000 cal of work. The net internal energy change of the system is ____.

35. The process of causing a non - spontaneous chemical reaction to take place by mean of electrical energy is called ___

Answer to supplementary exercise

Part I

1. True
2. True
3. False
4. True
5. False
6. True
7. False
8. False
9. True
10. True

Part II

11. D
12. A
13. C
14. B
15. C
16. A
17. B
18. B
19. C
20. C
21. B
22. A
23. C
24. D
25. C
26. A
27. D
28. A
29. A
30. B

Part III

31. Non - electrolyte
32. Electrodes
33. Anode reaction : $Zn \rightarrow Zn^{2+} + 2e^-$
   Cathode reaction : $Cu^{2+} + 2e^- \rightarrow Cu$
34. $\Delta E = -1000$ cal
35. Electrolysis

UNIT 10

Metals and Non - Metals

10.1 Introduction

Element is a pure substance that cannot be separated into simpler substances by chemical means. Based of property, elements are classified into three, metals, non - metals and metalloids.

Natural resources are resources that exist in nature. They are classified into two. These are:

1. Renewable natural resources
2. Non - renewable natural resources

1. Renewable Natural Resources

They are replaced by natural processes

Example
- ➥ Animals
- ➥ Plants
- ➥ Soil
- ➥ Fire wood
- ➥ Water

2. None - newable Natural Resources

They cannot be replaced by natural processes

Example
- ➥ Petroleum
- ➥ Minerals
- ➥ Nuclear energy
- ➥ Natural gas
- ➥ Coal

Metallic minerals are solid substances that contain one or more metallic elements

Example

Galena PbS

Cinnabar - HgS

Bauxite - $Al_2O_3 . 2H_2O$

Haematite - Fe_2O_3

Magnetite - Fe_3O_4

Malachite - $Cu(OH)_2 . CuCO_3$

Non - metallic minerals are minerals without metallic elements .

Example

Diamond, grave and mica

Occurence of Elements

☑ Most elements occur in combined state in the earth's crust and few elements occur in free state.

☑ 12 elements constitute 99.7 % by mass of the earth's crust. These elements in decreasing order are given below.

1. Oxygen
2. Silicon
3. Aluminum
4. Iron
5. Calcium
6. Sodium
7. Potassium
8. Magnesium
9. Hydrogen
10. Titanium
11. Chlorine
12. Phosphorus

Table 10 . 1 Abundance of Elements in the Atmosphere

Gas	Percentage by volume (%)
Nitrogen	78
Oxygen	21
Other gases	1

Illustrative Examples

1. Of the following substances, which one is NOT a natural resource ?

A. Metal ore B. Soil C. Glass D. Stone

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Solution

Glass is synthetic resource.

2. Which of the following natural resources is non - renewable ?

Answer : C

A. Petroleum B. Water C. Plants D. Animals

3. Which of the following natural resources is renewable ?

Answer : A

A. Iron ore
B. Coal
C. Water
D. Natural gas

4. Minerals are

Answer : C

A. Used as raw materials in chemical industries
B. Industrial products
C. Replenished with the same rate as they are consumed
D. Renewable resources.

Solution

Minerals are non - renewable natural resources used as raw materials in chemical industries.

Answer: A

5. The four most abundant elements in the earth's crust in decreasing order of abundances are _____

Solution

Oxygen > Silicon > Aluminum > iron

6. List some metals that occur in free state

Solution

Silver	Platinum	Ruthenium
Gold	Palladium	Osmium

7. The most abundant metal in the earth's crust is_____.

Answer : Aluminum

8. The most abundant nobel gas in the atmosphere is ___

Answer : Argon

10.2 General Properties of Metals and Production of Some Metals

General Properties of Metals

(i) Physical Properties of Metals

- ☑ All metals are solid at room temperature except **mercury** which is liquid.
- ☑ They are malleable (hammered into sheet)
- ☑ They are ductile (drawn into wire)
- ☑ They are luster (shiny)
- ☑ They are sonorous (produce metallic sound)
- ☑ They have generally high melting and boiling points. Their melting point ranges from - 39^{0}C (Hg) to 3410^{0}C (W).
- ☑ They are hard and have high tensil strength except alkali metals (Li, Na, K).
- ☑ They have high density except alkali metals
- ☑ They are good conductors of heat and electricity.

(ii) Chemical Properties of Metals

- ☑ Metals are electro positive elements because they have a tendency to lose electrons and form a positive ion.
- ☑ They are reducing agent (oxidize substance)
- ☑ They have 1, 2, or 3 valence electrons.
- ☑ They form basic oxides by reacting with oxygen. Few metals form amphoteric and acidic oxides.
- ☑ They react with dilute acids to form salt and hydrogen gas.
- ☑ They form salts by reacting with non - metals.

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Table 10 . 2 Reactivity Order

Reactivity series	Reation with dilute acids	Reation with air (oxygen)	Reation with water	Ease of extraction
Potassium(K)	Produce H_2 with decreasing vigour	burn very brightly and vigorously	produce H_2 with decreasing vigour with cold water	difficult to extract
Sodium(Na)				
Calcium(Ca)		burn to form an oxide with decreasing vigour	react with steam with decreasing vigour	easier to extract
magnessium (Mg)				
Aluminum (Al)				
zinc(Zn)				
Iron(Fe)				
Tin(Sn)				
Lead(Pb)				
Hydrogen(H)	do not react with dilute acids	react slowly to form oxide	do not react with cold water or steam	
(non- metal)				
copper (Cu)		do not react		found as free element
mercury (Hg)				
silver (Ag)				
Gold(Au)				

Extraction of Metals

Metallurgy is the science and technology of extracting metals from their natural sources and preparing them for practical use

There are four steps in metallurgy. These are

1. **Concentrating ore** - is increasing the amount of metal containing ore . This can be done by
   - ➥ Gravity separation
   - ➥ Magnetic separation
   - ➥ Froth flotation
   - ➥ Amalgamation
   - ➥ Leaching
2. **Pre treatment** is converting an ore to form that will be easy to reduce.

3. Production of metals.
   - ➥ Metals are produced by **reduction process.**

There are two types of reductions. These are

(i) **Chemical reduction** is manufacturing metals using reducing agents.

Example

Less active metals such copper, zinc, chromium, tin, tungsten etc.

(ii) **Electrolytic reduction** is manufacturing metals using electrolysis.

Example

More active metals such as sodium, magnesium, calcium and aluminum.

4. Refining and Alloying

Refining is purification of metals.

Alloys are homogeneous mixtures of metals which are stronger and corrsion resistance than the parent metals.

Table 10 . 3 Composition and Uses of Some Alloys

Alloy	Composition	Use
Brass	65 % copper, 35 % zinc	Jewellery, machine bearing, electrical connections, door
Bronze	90 % copper, 10 % tin	Castings, machine parts
cupro - nicekel	75 % copper, 25 % nickel	Coinage metal
Cupro - nickel	30 % copper, 70 % nickel	Turbine blades
Duralumin	95 % aluminum, 4 % copper, 1 %, Mn	Aircraft construction, bicycle parts
Magnalium	70 % aluminum, 30 % magnesium	Aircraft construction
Pewter	30 % lead, 70% tin, a small amount of antimony	Plates, ornaments and drinking mugs
Solder	70 % lead, 30 % tin	Connecting electrical wiring

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Nichrome	80 % nickel, 20 % chromium	Flame test of ions.

A carat is equal to the mass unit of gold in 24 mass units of the alloy.

- 24 carat represents pure gold
- 21 carat gold contains 21 parts of pure gold by mass in 24 carats.
- 18 carat gold contains 18 parts of pure gold by mass in 24 carats.
- 14 carat gold contains 14 parts of pure gold by mass in 24 carats.

Example

Calculate the mass percent of gold in 18 carat gold.

Solution

$$\text{Mass \% gold} = \frac{\text{cerat of gold}}{\text{24 carat}} \times 100 = \frac{18}{24} \times 100 = 75\ \%$$

Thus, in 18 carat gold, 75% is gold and 25 % is copper or siliver

Illustrative Examples

9. All of the following are true about metals, **EXCEPT**

A. They are good conductors of heat and electricity
B. They are hard and brittle
C. They are malleable and ductile
D. They are solids except mercury.

Answer: B

10. The most reactive metal in the reactivity series is

A. Gold B. Zinc C. Potassium D. Sodium

Solution

In the reactivity series, the most active metal is potassium and the least active metal is gold. Answer: C

11. All of the following processes are used to concentrate ore. **EXCEPT**

A. Gravity separation
B. Magnetic separation
C. Froth flotation
D. Alloying

Solution

Alloying is important to increase strength and reduce corrosion. Answer: D

12. Explain the difference between roasting and calcination.

Solution

Roasting is heating ore in the presence of air to oxide but calcination is heating an ore in the absence of air.

13. Write the first four reactive metals in order of increasing reactivity.

Solution

Magnesium < calcium < sodium < potassium

14. List four metals that are less reactive than hydrogen.

Solution

Copper, mercury, silver and gold.

Occurrence, Extraction, property and uses of some metals

1. Aluminum (Al)

Occurrence

- ☑ It does not exist free in nature
- ☑ Its main ore is bauxite ($Al_2O_3 . 2H_2O$)
- ☑ Other minerals containing alminum are :
  - Orthoclase - $KAlSi_3O_6$
  - Beryl - $Be_3Al_2Si_6O_{18}$
  - Cryolite - Na_3AlF_6
  - Corundum - Al_2O_3
  - China clay - $Al_2Si_2O_7 . 2H_2O$

Extraction

- ☑ **Aluminum** is industrially manufactured by **Hall process** using electrolysis from bauxite
- ☑ **Cryolite** is added to bauxite to lower its melting point.

Cathode reaction : $4Al^{3+} + 12e^- \rightarrow 4Al(l)$

Anode reaction : $6O^{2-} \rightarrow 3O_2 + 12e^-$

$3C(s) + 3O_2(g) \rightarrow 3O_2(g)$

Overall reaction : $4Al^{3-} + 6O^{2-} + 3C \rightarrow 4A(l) + 3CO_2(g)$

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Physical Properties of Aluminum

- It is a soft, white metal.
- It is a good conductor of heat and electricity
- It has a brilliant luster.
- It is malleable and ductile

Chemical Properties of Aluminum

(i) Aluminum reacts with non - metals.

Example

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$
$$4Al(s) + 6S(s) \rightarrow 2Al_2S_3(s)$$
$$2Al(s) + N_2(g) \rightarrow 2AlN(s)$$
$$2Al(s) + 3F_2(g) \rightarrow 2AlF_3(s)$$

(ii) Aluminum is amphoteric metal, reacts with acids and bases.

$$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(s) + 3H_2(g)$$
$$2Al(s) + 2NaOH(aq) + 6H_2O(l) \rightarrow 2NaAl(OH)_4(aq) + 3H_2(g)$$

Aluminum is passive towards nitric acid due to a thin layer transparent film of Al_2O_3 formed when aluminum comes into contact with air.

(iii) Aluminum displaces iron from iron (III) oxide

This reaction is called thermite welding process

$$2Al(s) + Fe_2O_3(s) \rightarrow 2Fe(l) + Al_2O_3(s)$$

The mixture of powdered aluminum and iron (111) oxide is called **thermite.**

(iv) Aluminum reacts with chromium (III) oxide

$$Al(s) + Cr_2O_3(s) \rightarrow 2Cr(l) + Al_2O_3(s)$$

Uses of Aluminum

- ☑ Its alloys, duralumin (Al/Mg/Cu) and magnalium (Al/Mg) are light, strong and corrosion resistant.
- ☑ It is used to make cooking utensils, brewing vats, food procelling vessels and packing and copping milk bottles.
- ☑ It is used for making electrical conductors for transmission lines.
- ☑ The very high amount of heat released by the thermite reaction makes it useful in welding large steel parts such as propeller shafts and ladder posts on ships and heavy machinery.
- ☑ It is used as reducing agent for small scale extraction of chromium.

Illustrative Examples

15. Hall's process of production of aluminum involves

A. Heating cryolite with CO
B. Electrolytic reduction of Al_2O_3
C. Displace Al from its salts using active metals
D. Reducing $AlCl_3$ using coke.

Solution

Aluminum is manufatured industrially by electrolytic reduction of Al_2O_3.

Answer : B

16. Which property of aluminum makes it preferably best for construction of air planes frame work ?

A. High ductility
B. Good electrical conductivity
C. Resistance to corrosion
D. Lightness and strength

Solution

Uses of aluminum	Reason
➥ Construction of air planes	➥ Lightness and strength
➥ Making cooking utensils	➥ Good conductivity of heat
➥ Electrical cables	➥ Light, ductile, corrosion resistance

Answer : D

17. Which of the following elements is the third most abundant element in the earth's crust ?

A. Silicon
B. Aluminum
C. Calcium
D. Iron

Answer : B

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18. What is the purpose of cryolite in the production of aluminum by Hall process ?

Solution

It is used to lower the melting point of bauxite.

19. List the most important alloys of aluminum.

Solution

Duralumin = Al + Mg + Cu

mangalium = Al + Mg

Aluminum bronze = Al + Cu

20. What is thermite reaction ?

Solution

Thermite reaction is the reaction between aluminum and iron (III) oxide.

$2Al + Fe_2O_3 \rightarrow 2Fe + Al_2O_3$, $\Delta H = -84\ kJ/mol$

2. Iron (F e)

Occurrence

- It does not exist free in nature
- Its main ore is hematite (Fe_2O_3)

Other minerals of iron are :

- Limonite - $Fe_2O_3 . H_2O$
- Magnetite - Fe_3O_4
- Siderite - $FeCO_3$
- Iron pyrite or fool's gold - FeS_2

Extraction

Iron is industrially manufactured in blast furnace.

Raw materials

- Iron ore - source of iron
- Lime stone - remove impurities
- Hot air - source of heat
- Coke - reducing agent

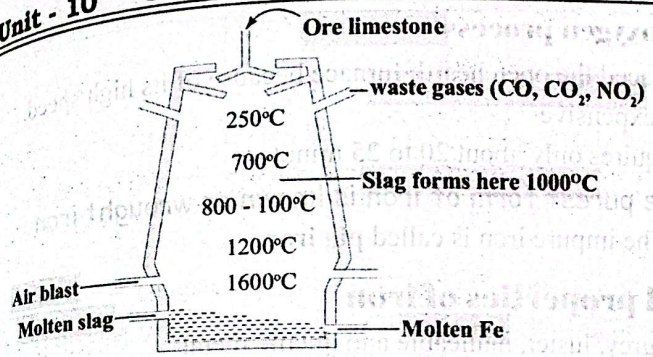

Fig 10. 1 Extraction of Iron

Chemical Reduction of Hematite is Shown Below

1. $C + O_2 \rightarrow CO_2 + heat$
2. $CO_2 + C \rightarrow 2CO$
3. $2Fe_2O_3 + CO \rightarrow 2Fe_3O_4 + CO_2$

   $Fe_3O_4 + CO \rightarrow 3FeO + CO_2$

   $FeO + CO \rightarrow Fe + CO_2$
4. $CaCO_3 \rightarrow CaO + CO_2$
5. $CaO + SiO_2 \rightarrow CaSiO_3$

- The iron obtained from blast furnace is called **pig iron.**
- **Pig iron** is converted into steel by three processes.

1. The Bessemer Converter

- This process requires only 15 minutes
- The quality of steel produced by this process is variable
- The reaction is difficult to control
- Uses $CaCO_3$ and $MgCO_3$

2. The open - hearth furnace

- Lined with MgO and CaO
- The process requires 8 -10 hours
- More easily controlled reaction
- High quality steel is formed

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3. Basic oxygen process

- Replaced the open hearth furnace because of its high speed.
- It is expensive
- It requires only about 20 to 25 minutes

Note: The purest form of iron is known as wrought iron

- The impure iron is called **pig iron.**

Physical properties of iron

- It is grey, luster, malleable and ductile metal
- It is good conductor of heat and electricity
- It is heavy metal with density 7.87 g/cm^3.

Chemical properties of iron

Iron react with air

$$4Fe + 3O_2 \rightarrow 2Fe_2O_3$$

It reacts with acids

$$Fe + 2HCl \rightarrow FeCl_2 + H_2$$

$$2Fe + 6\,H_2SO_4 \rightarrow Fe_2(SO_4)_3 + 3SO_2 + 6H_2O$$

It **reacts** with non -metals such as sulphur and chlorine.

$$Fe + S \rightarrow FeS$$

$$2Fe + 3Cl_2 \rightarrow 2FeCl_3$$

Uses of Iron

☑ **Iron** is the cheapest metal and in the form of steel, the most useful. There are two types of steels, **carbon steel and alloy steel.**

☑ **Carbon steel** can be classified as mild steel, medium steel or high carbon steel.

Table 10 . 4 The Three Types of Carbon Steel

Types of steel	Percentage of carbon	Uses
Mild steel	≤ 0.2	General engineering, motor car bodies, wire, piping nuts and bolts
Medium steel	0.3 - 0.6	Beams and girders, springs
High carbon steel	0.6 - 1.5	Drill bits, knives, hommers chisels.

Alloy steel is also classified as **stain less steel and tool steel.**

- Stain less steel = Fe + Cr + Ni
- **Tool steel** = Tungsten and manganese steel
- Alloy steel is used in rock drills, cutting edges and parts of machinery.

Illustrative Examples

21. What is the role of calcium carbonate in the blast furnace during the extraction of iron ?

A. Used as a raw material
B. To facilitate the removal of impurities
C. To reduce iron ore to iron
D. To oxidize carbon to carbon monoxide

Solution

Calcium carbonate is used to remove impurities in the form of slag.

Answer : B

22. Among the following , identify the purest form of iron.

A. Wrought iron
B. Cast iron
C. Steel
D. Pig iron

Answer : A

23. What is the main ore of iron ?

Answer : Hematite

24. Complete and balance the following

a. $Fe + O_2 \rightarrow$
b. $Fe + HCl \rightarrow$
c. $Fe + Cl_2 \rightarrow$
d. $Fe + S \rightarrow$
e. $Fe + H_2SO_4 \rightarrow$

Solution

a. $4Fe + 3O_2 \rightarrow 2Fe_2O_3$

b. $Fe + 2HCl \rightarrow FeCl_2 + H_2$

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c. $Fe + 3Cl_2 \rightarrow 2FeCl_3$

d. $Fe + S \rightarrow FeS$

e. $Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$

25. What are the two common oxidation number of iron ?

Solution

- Iron (11) and Iron (111)
- Iron (I I) compounds have pale - green color and iron (I I I) compounds have yellowish brown color.

26. What are the component metals for stainless steel ?

Answer : Iron, Chromium and nickel

27. Pig iron is much cheaper than steel but is not used in construction. Explain why pig iron is not suitable for erecting buildings and bridges.

Explanation

The carbon content of pig iron is too high, it makes pig iron relatively hard and brittle.

28. What is the purpose of coke in the blast furnace ?

Solution

It is reducing agent

29. Write the balanced acid - base reaction that occurs in the blast finance.

Solution

$CaO + SiO_2 \rightarrow CaSiO_3$

30. Write the components of pig iron

Solution

Pig iron contains about 4 % carbon, 2 % silicon, a trace of sulphur and 1 % of phosphorus and manganese. Sulphur is probably the worst impurity and must be avoided.

3. Copper (Cu)

Occurrence

- ☑ Copper exists both in free and combined states
- ☑ Its main ore is **chalcopyrite ($CuFeS_2$)**
- ☑ Other minerals of copper are :

- Chalcocite ----- Cu_2S
- Covellite ----- CuS
- Cuprite ----- Cu_2O
- Malachite ---- $(CuCO_3 . Cu(OH)_2)$
- Tenorite ----- CuO

Extraction

It is produced by **chemical reduction**

steps

1. Concentrating ore by floatation

$2CuFeS_2 + 4O_2 \rightarrow Cu_2S + 2FeO + 3SO_2$

Limestone removes iron (11) oxide.

$FeO + SiO_2 \rightarrow Fe\ SiO_3$

2. Smelting or reduction

$Cu_2S\ (s) + O_2\ (g) \rightarrow 2\ Cu(s) + SO_2\ (g)$

3. Electrorefing /purification / using electrolysis.

Physical Properties of Copper

- It is malleable and ductile
- It is very good conductor of heat and electricity
- It is heavy metal with a density of 8.96 g / cm^3
- It has high melting and boiling points.

Chemical Properties of Copper

Copper cannot displace hydrogen from dilute acids, however, it can be oxidized by oxidizing acids such as dilute and concentrated sulphuric acid.

$3Cu\ (s) + 8HNO_3\ (dilute) \rightarrow 3\ Cu\ (NO_3)_2\ (aq) + 2NO\ (g) + 4H_2O\ (1)$

$Cu\ (s) + 2H_2SO_4\ (conc) \rightarrow CuSO_4\ (aq) + SO_2\ (g) + 2H_2O\ (1)$

- Copper corrodes in moist air.

$2\ Cu\ (s) + H_2O\ (1) + O_2\ (g) + CO_2\ (g) \rightarrow Cu_2\ (OH)_2\ CO_3\ (s)$ basic copper carbonate

Uses of copper

- ☑ Copper is widely used in the electrical industry
- ☑ It is also used to make alloys such as brass (Cu + Zn),

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bronze (Cu + Sn) and copper - nickel (Cu + Ni)

- ☑ Brass is used for hard ware taps, terminals and pipes.
- ☑ Bronze is used to make coins, medals, bells and machinery parts.
- ☑ Copper compounds such as $CuCl_2$, $CuCO_2$, $CuCO_3$ and $Cu(OH)_2$ are used as pesticides.

Illustrative Examples

31. Which of the following is the correct sequence for the industrial production of pure copper from chalcopyrite ($CuFeS_2$) ?

A. Roasting of the ore → smelting of the ore in limestone → concentrating the ore → reducing the ore in limited oxygen → electrolytic refining.

B. Concentrating the ore → smelting of the ore in lime stone → roasting of the ore in oxygen → reducing the ore in limited oxygen → electrolytic refining.

C. Smelting of the ore → roasting of the ore in oxygen → reducing the ore in limited oxygen → electrolytic refining

D. Concentrating the ore → roasting of the ore in oxygen → smelting of the ore in limited oxygen → electrolytic refining

Answer : D

32. Which one of the following elements is the component of both brass and bronze ?

A. Ni B. Zn C. Cu D. Sn

Solution

Bronze = Cu + Sn, Brass = Cu + Zn,

Answer : C

33. Which of the following is the principal ore of copper ?

A. $Cu FeS_2$ B. Cu_2S C. CuO D. $CuCO_3$

Answer : A

34. Write the disproportionation reaction of the unstable Cu^+.

Answer : $2Cu^+(aq) \rightarrow Cu^{2+}(aq) + Cu(s)$

35. What is the color of copper (I) and copper (11) compounds ?

Solution

Copper (I) compounds are colorless while copper (II) compounds are blue in color.

36. What is the metal that turns green due to the formation of verdigris ?

Answer : Copper

4. Sodium (Na)

Occurrence

- Main ore sodium chloride (NaCl)

Other minerals of sodium are:

- Borax ($Na_2B_4O_7 . 10H_2O$)
- Trona ($Na_2CO_3 . NaHCO_3 . 2H_2O$)
- Chile salt peter ($NaNO_3$)

Extraction

Na is produced by **down process** using 40 % NaCl and 60 % $CaCl_2$ in fused state.

$CaCl_2$ is used to lower the melting point of NaCl from 1080K to about 850K

Cathode reaction : $2Na^+ + 2e^- \rightarrow 2Na(l)$

Anode reaction : $2Cl^- \rightarrow Cl_2(g) + 2e^-$

Overall reaction : $2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$

Chemical properties of sodium

It reacts with :

a. Water : $2Na + 2H_2O \rightarrow 2NaOH + H_2$

b. Air : $6Na + 2O_2 \rightarrow Na_2O_2 + 2Na_2O$

c. Halogen : $2Na + Cl_2 \rightarrow 2NaCl$

d. Hydrogen : $2Na + H_2 \rightarrow 2NaH$

Uses of sodium

- ☑ Sodium metal is a strong reducing agent and this accounts for its major uses.
- ☑ It is used to obtain metals such as titanium and zirconium.
- ☑ It is used in sodium lights for lighting high ways.

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☑ It is also used to cool fuels in nuclear reactors

☑ A sodium lead alloy is used to make tetraethyl lead which is added to petrol to act as an anti knock agent.

5. Calcium

Occurrence

Its main ore is calcium chloride ($CaCl_2$).

Other minerals of calcium are

- Limestone - $CaCO_3$
- Dolomite - $Mg\,CO_3 . CaCO_3$

Gypsum - $CaSO_4 . 2H_2O$

Anhydrite - $CaSO_4$

Rock phosphate - $Ca_3(PO_4)_2$

Extraction

By electrolysis of molten $CaCl_2$

Anode reaction : $2Cl^- \rightarrow Cl_2 + 2e$

Cathode reaction : $Ca^{2+} + 2e^- \rightarrow Ca\,(1)$

Over all reaction : $Ca^{2+}(aq) + 2Cl^-(aq) \rightarrow Ca\,(1) + Cl_2\,(g)$

Chemical properties of calcium

Calcium reacts with :

a. Water : $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$

b. Air : $2Ca + O_2 \rightarrow 2CaO$

c. Halogens : $3Ca + X_2 \rightarrow CaX_2$ (X - halogen)

d. Nitrogen : $3Ca + N_2 \rightarrow Ca_3N_2$

e. Hydrogen : $Ca + H_2 \rightarrow CaH_2$

Uses of Calcium

☑ It is major components of teeth and bones .

☑ It plays an important role in heart actions, blood clotting, muscle contraction and nerve transmission.

☑ It is used as reducing agent in the production of certain metals such as Cr, Zr. Th and U.

☑ Calcium is useful as deoxidizer in steel production.

☑ It is also used to remove gases (N_2, O_2) from electronic tubes.

Illustrative Examples

37. Which of the following metal is used to cool the fuel in nuclear reactors ?

A. Na B. Ca C. Li D. Mg

Answer : A

38. All of the following are minerals of calcium **EXCEPT.**

A. Limestone B. Dolomite C. Trona D. Gypsum

Answer : C

39. What is the purpose of calcium chloride in the production of sodium by downs process ?

Solution

It is used to lower the melting point of NaCl

40. Write the two important methods used to extract calcium.

Solution

- Electrolysis of molten Ca Cl_2

$$CaCl_2 \xrightarrow{\text{electrolysis}} Ca(1) + Cl_2\,(g)$$

- Reduction of calcium oxide with aluminum in a vacuum, where the calcium produced distills off.

$$3CaO(s) \xrightarrow{\text{electrolysis}} Ca\,(1) + Cl_2(g)$$

41. Sodium is used in sodium lights for lighting high ways. Explain why ?

Solution

Because its yellow light penetrates fog well.

10.3 Production of some Important non - metals

1. Nitrogen

Occurrence

Note : The general properties of non - metals are the opposite

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of that of metals.

- ☑ Nitrogen occurs free in the atmosphere as simple gaseous diatomic molecules, N_2.
- ☑ It also occurs in plants and animals in the form of proteins.
- ☑ The most important mineral sources of nitrogen are salt peter (KNO_3) and chile salt peter ($NaNO_3$)

Physical Properties of Nitrogen

- ➥ It is a colorless, odourless and tasteless gas
- ➥ It is slightly less dense than air
- ➥ It boils at - 196 °C

Chemical Properties of Nitrogen

Nitrogen is less reactive non -metal due to the nitrogen - nitrogen triple bond.

It reacts with active metals

Example

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

$$N_2(g) + O_2(g) \rightarrow 2NO(g)$$

Preparation of Nitrogen

Industrially, nitrogen is manufatured by **fractional distillation of liquid air**

In the laboratory, nitrogen is prepared by thermal decomposition of ammonium nitrite :

$$NH_4NO_2(s) \rightarrow 2H_2O(1) + N_2(g)$$

Uses of nitrogen

- ☑ It is used to creat inert atmosphere.
- ☑ Liquid nitrogen is used as a refrigerant for preserving blood bull's semen and some foods.
- ☑ The most important use of nitrogen is in the production of ammonia, NH_3

Illustrative Examples

42. The most abundant gas in the atmosphere is

A. N_2 B. O_2 C. CO_2 D. Ar

Solution

78 % by volume of the air is nitrogen. **Answer : A**

43. Which of the following does not describe the chemical properties of nitrogen ?

A. It is a relatively reactive gas
B. It combines with hydrogen to form ammonia
C. Reacts with group IA and IIA
D. It reacts with oxygen **Answer : A**

44. Which of the following is not the uses of nitrogen ?

A. Creating inert atmosphere
B. For preserving blood and some foods
C. Micro nutrient for plants
D. Making amino acids **Answer : C**

45. It is difficult to imagine a living system that does not contain nitrogen. Why ?

Solution

The reason for this is that nitrogen is an essential component of proteins, nucleic acids, vitamins and hormones that make life possible.

46. Nitrogen is less reactive element, Why ?

Solution

It is due to the nitrogen - nitrogen triple bond ($N \equiv N$)

2. Phosphorus (P)

Occurrence

It does not found free in nature

It occurs in combined state as :

- ➥ Calcium phosphate - $Ca_3(PO_4)_2$
- ➥ Fluorapatite - $Ca_5(PO_4)_3F$

Its compounds are found in living things.

Physical properties of phosphorus

Phosphorus has two allotropes. White and red phosphorus.

Table 10 . 5 Different Properties of White and Red Phosphorus

White phosphorus	Red phosphorus
➥ Boils at 287°C	➥ Sublimes at 416 °C
➥ Density 1.8 g/cm³	➥ Density 2.16 g/ml
➥ White waxy solid	➥ Dark red power
➥ Glows in dark	➥ Does not glow in the dark
➥ Soluble in CS_2	➥ Not soluble in CS_2
➥ Stored in water	➥ Stored in CS_2
➥ Extremery poisonous	➥ Relatively non - toxic compared to white phosphorus.

Chemical Properties of Phosphorus

It reacts with metals

Example

$$6Ca(s) + P_4(s) \rightarrow 2Ca_3P_2(s)$$

It react with non - metals.

$$P_4(s) + 3O_2(g) \rightarrow P_4O_6(s)$$

$$P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$$

$$4P_4(s) + 5S_8(s) \rightarrow 4P_4S_{10}(s)$$

$$8P_4(s) + 3S_8(s) \rightarrow 8P_4S_3(s)$$

$$P_4(s) + 6Cl_2(g) \rightarrow 4PCl_3(s)$$

$$P_4(s) + 10Cl_2(g) \rightarrow 4PCl_5(s)$$

Preparation of phosphorus

White phosphorus can be obtained by heating calcium phosphate with coke and silica/ s and / at very high temperature.

$$2Ca_3(PO_4)_2 + 6SiO_2 + 10C \rightarrow 6CaSiO_3 + P_4 + 10CO$$

Red Phosphorus is prepared by heating white phosphorus in the absence of air at about 300°C for several days.

$$nP_4 \rightarrow (P_4)n$$

White Phosphorus Red Phosphorus

Uses of Phosphorus

- ☑ Phosphorus is used to make matches.
- ☑ White phosphorus is used to make phospheric acid.
- ☑ It is used to make fireworks, smoke bombs, rat poisons and tracer bullets.
- ☑ It is essential for plant growth and to make fertilizers.

Illustrative Examples

47. Which of the following elements, occurs in nature only in combined form ?

A. Nitrogen C. Sulphur
B. Phosphorus D. Carbon

Answer : B

48. Which of the following is not the purpose of phosphorus ?

A. It is used to make fertilizer
B. It is used to make matches
C. It is used to make sulphuric acid
D. It is used to make rat poison

Answer : C

49. What are the two important allotropes of phosphorus ? Which one is more stable ?

Solution

White and red phosphorus. Red phosphorus is more stable than white phosphorus.

50. Define allotropes.

Solution

Allotropes are different forms of the same element with the same physical state and isotopic composition but different physical properties, structure and reactivity.

51. Complete and balance the following equation

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a. $P_4 + Li \rightarrow$ b. $Ca + P_4 \rightarrow$

Solution

a. $P_4 + 12\ Li \rightarrow 4\ Li_3P$ b. $6Ca + P_4 \rightarrow 2Ca_3P_2$

3. Oxygen

Occurrence

☑ Oxygen is the most abundant element on earth.

☑ The earth's crust is 46.6 % oxygen by mass, the atmosphere contains 21 % molecular oxygen by volume and the oceans contain 86 % oxygen by mass

Preparation of oxygen

☑ Industrially, oxygen is manufactured by fractional **distillation of liquid air.**

☑ In the laboratory oxygen can be prepared by the following methods.

1. By decomposition of hydrogen peroxide

$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$

2. By decomposition of potassium chlorate

$2K\ ClO_3\ (s) \rightarrow 2KCl\ (s) + 3O_2\ (g)$

3. By decomposition of mercury (II) oxide

$2HgO\ (s) \rightarrow 2Hg\ (l) + O_2(g)$

Physical properties of oxygen

➥ **Oxygen** is colorless, odorless and tasteless gas at room temperature.

➥ It has two allotropes, these are oxygen (O_2) and ozone (O_3)

Chemical properties of Oxygen

Oxygen reacts with almost all elements except noble gases and inert metals such as gold, platinum palladium etc.

Example

$2Ca + O_2 \rightarrow 2CaO$

$2C + O_2\ (\text{limited}) \rightarrow 2CO$

$C + O_2\ (\text{excess}) \rightarrow CO_2$

$S_8 + 8O_2 \rightarrow 8SO_2$

$P_4 + 5O_2\ (g) \rightarrow P_4O_{10}\ (s)$ ---excess oxygen

$P_4 + 3O_2\ (g) \rightarrow P_4O_6\ (s)$ ------ limited oxygen

Uses of Oxygen

➥ Oxygen is widely used as an oxidizing agent.

➥ Most of the oxygen produced is use in the production of steel.

➥ It is used in the production of metals

➥ With acetylene, oxygen is used in oxyacetylene welding.

➥ It is used to support life in medicine, in air and space travel and in submarines.

➥ It is also used as bleaching agent for compounds and in biological treatment of waste water etc.

4. Sulphur Occurrence

➥ Sulphur is found as the elemental form in nature

➥ It is also found in combined form as :

Sulphide ore

➥ Galena - PbS

➥ Cinnabar - HgS

➥ Pyrite (fool's gold) - FeS_2

➥ Sphalerite or zinc bland - ZnS

Sulphates such as :

➥ Gypsum - $CaSO_4 \cdot 2H_2O$

➥ Barite - $BaSO_4$

➥ Natural gas contain hydrogen sulphide (H_2S)

Production of Sulphur

Sulphur is extracted from underground deposits of elemental sulphur by the **Frasch process**

➥ **Super heated steam** - melts the sulphur in the ground

➥ **Compressed air** - forces the forth of sulphur from the ground to up

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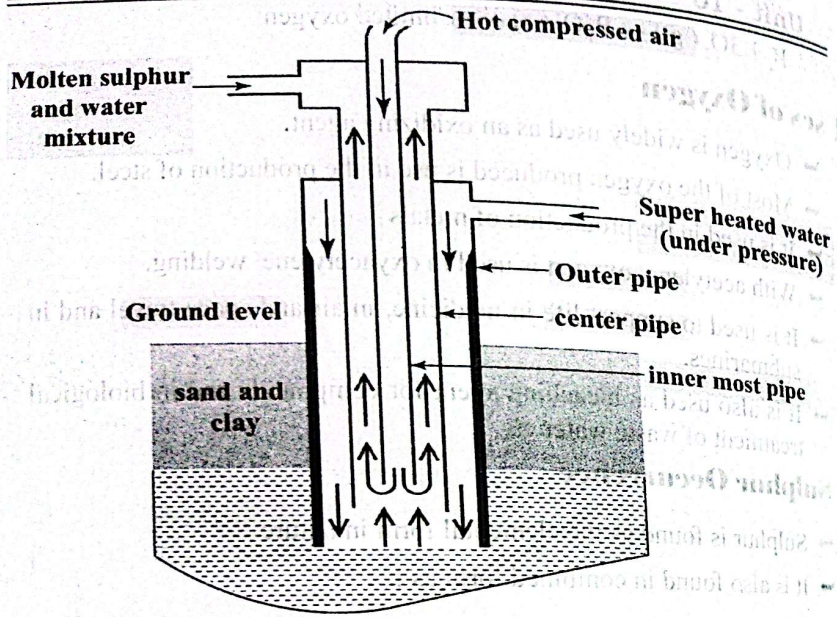

Fig 10. 2 The Frasch process

Physical properties of Sulphur

Sulphur has two important allotropes. These are **rhombic and monoclinic sulphur**

Table 10 . 6 Different Properties of Monoclinic and Rhombic Sulphur.

Rhombic Sulphur	Monoclinic Sulphur
➥ Yellow, tasteless and oderless solid ➥ Stable form of sulphur containing S_8 molecules ➥ Insoluble in water ➥ Soluble in CS_2 ➥ Converted to monoclinic sulphur when heated ➥ Stable below 96^0c	➥ Consists of S_8 molecules ➥ Less stable form of sulphur ➥ Insoluble in water but soluble in CS_2 ➥ Stable between 96^0C

Chemical properties of Sulphur

Sulphur reacts with active metals

Example

$$16\,K(s) + S_8(s) \rightarrow 8K_2S(s)$$

Sulphur reacts with non - metals.

Example

$$S_8(s) + 8O_2(g) \rightarrow 8SO_2(g)$$
$$S_8(s) + 16\,F_2(g) \rightarrow 8SF_4(g)$$
$$S_8(s) + 24\,F_2(g) \rightarrow 8SF_6(g)$$

Uses of Sulphur

☑ Sulphur is used widely in making sulphuric acid
☑ It is used in the vulcanization of rubber
☑ Sulphur is used in the manufacture of skin ointments and drugs.
☑ It is a good insulator

Illustrative Examples

52. Which of the following is used in the production of steel ?

A. Nitrogen
B. Sulphur
C. Oxygen
D. Phosphorus

Solution

Oxygen is widely used in the production of steel.

Answer : C

53. Which of the following processes is used for the industrially production of oxygen ?

A. Fractional distillation of liquid air
B. Photosynthesis
C. Thermal decomposition of oxides
D. Electrolysis of water

Answer : A

54. All of the following compounds can be used as raw material to manufacture oxygen in the laboratory. **EXCEPT**

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A. $KClO_3$ B. HgO C. Ag_2O D. $CaCO_3$

Answer :D

55. Which of the following elements is used to vulcanize rubber ?

A. Phosphorus C. Carbon
B. Sulphur D. Chlorine

Answer : B

56. Which one of the following processes is used to obtain sulphur from the underground sulphur deposited ?

A. Frasch process C. Solvay process
B. Ostwald process D. Contact process

Solution

Ostwald process - Nitric acid
Solvay process - Sodium carbonate
Contact process - Sulphuric acid

Answer : A

57. What are the common oxidation states of oxygen in compounds ?

Solution

- - 2 in most compounds
- - 1 in peroxides
- - 1 / 2 super oxides
- + 2 in OF_2

58. What are the two allotropes of oxygen ?

Solution

O_2 and O_3

59. What are the two gases manufactured industrially by the same method ?

Solution

N_2 and O_2 by fractional distillation.

60. What is the main purpose of sulphur ?

Solution

To produce sulphuric acid.

61. What are the two important allotropes of sulphur ?

Solution

Monoclinic and rhombic sulphur.

5. Chlorine

Occurrence

- Chlorine is found in nature in combined form as chlorides of sodium, potassium, calcium and magnesium.
- Sodium chloride is the main ore of chlorine.

Preparation of chlorine

Industrially, chlorine is manufactured by electrolysis of **concentrated sodium chloride solution / brine /.**

$$2NaCl(aq) + 2H_2O(l) \rightarrow Cl_2(g) + 2NaOH(aq) + H_2(g)$$

In the laboratory chlorine can be prepared by the following reaction :

$$MnO_2(s) + 4HCl(aq) \rightarrow MnCl_2(s) + Cl_2(g) + 2H_2O(l)$$

Physical properties of Chlorine

- It is greenish - yellow gas
- It metals at - 101°C and boils at - 34°C.
- It is poisonous gas

Chemical Properties of Chlorine

It reacts with active metals

Example

$$2Fe + 3Cl_2 \rightarrow 2FeCl_3$$
$$Zn + Cl_2 \rightarrow ZnCl_2$$

It reacts with hydrogen

Example

$$Cl_2 + H_2 \rightarrow 2HCl$$

Chlorine displaces bromine and iodine from their compounds.

Example

$$Cl_2(g) + 2\,NaBr(s) \rightarrow Br_2(g) + 2\,NaCl(s)$$
$$Cl_2(g) + 2NaI(s) \rightarrow I_2(g) + 2NaCl(s)$$

Chlorine reacts with water to form hydrochloric acid and hypochlorous

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acid. This reaction of chlorine is called **disproportionation reaction.**

$Cl_2(g) + H_2O(l) \rightarrow HCl + HClO$

Uses of Chlorine

- ☑ Chlorine is a very strong oxidizing agent,which is used commercially as a bleaching agent and as disinfectant
- ☑ Chlorine is widely used for producing safe drinking water.
- ☑ It is also extensivery used in the production of dyestuffs, textiles, petroleum products,medicines, antiseptics, insecticides, plastics.
- ☑ It is also used to make solvents such as CCl_4, chloroform ($CHCl_3$), Dichlorodichloro methylene (CH_2Cl_2)

Illustrative Example

62. Which one of the following chemicals is used to disinfect water ?

A. Fluorine B. Nitrogen C. Oxygen D. Chlorine

Answer : D

63. Which of the following is the most abundant halogen ?

A. Cl_2 B. F_2 C. I_2 D. Br_2

Answer : A

64. Why chlorine displaces only bromine and iodine but not fluorine from their compounds ?

Solution

Because chlorine is less reactive than fluorine.

65. Balance and complete the following reactions

a. $F_2 + KCl \rightarrow$

b. $Cl_2 + KBr \rightarrow$

c. $Br_2 + KI \rightarrow$

d. $Cl_2 + KI \rightarrow$

Solution

a. $F_2 + 2\,KI \rightarrow 2\,KF + Cl_2$

b. $Cl_2 + 2\,KBr \rightarrow 2KCl + Br_2$

c. $Br_2 + 2\,KI \rightarrow 2\,KBr + I_2$

d. $Cl_2 + 2\,KI \rightarrow 2\,KCl + I_2$

Solved problems

66. What are the three important methods that converts pig iron into steel ?

Solution

- The Bessemer converter
- The open hearth furnance
- The basic oxygen converter

67. Classify the following as heavy or light metals.

a. Au c. Al e. K g. Ag

b. Fe d. Cu f. Mg h. Ca

Solution

- a, b, d, and Ag are heavy metals
- c, e, f, and h are light metals.

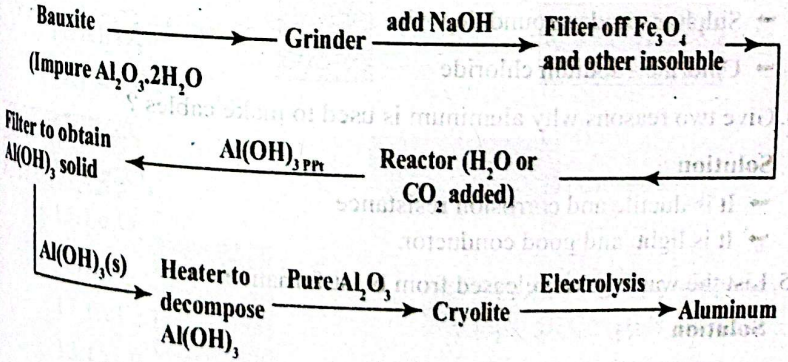
68. Write the steps in the extraction of aluminium by Hall process.

Solution

1. Purification of bauxite
2. Electrolysis of pure bauxite

69. Write summary of extraction of aluminium

Solution



70. Write the steps in the extraction of iron.

Solution

1. Preparation of the ore for reduction
2. Reduction of the oxides of iron

71. Write the steps in the extraction of copper.

Solution

1. Froth flotation for the concentration of the ore
2. Reduction of the concentrated ore
3. Purification of copper

72. What is the approximate formula of rust and greenish layer of copper.

Solution

Rust = $Fe_2O_3 . x H_2O$

Greenish layer = $Cu(OH)_2 . CuCO_3$

73. Write the main sources of nitrogen. Oxygen, phosphorus, sulphur and chlorine.

Solution

- Nitrogen - atmospheric air
- Oxygen - atmospheric air
- Phosphorus - calcium phosphate
- Sulphur - underground deposit
- Chlorine - sodium chloride

74. Give two reasons why aluminum is used to make cables ?

Solution

- It is ductile and corrosion resistance
- It is light and good conductor.

75. List the waste gases released from blast furnance.

Solution

CO, CO_2 and NO_2

Supplementary Exercise

I. True / False Items

1. Chlorine is commercially used as a bleaching agent and disinfectant.
2. Iron is the second most abundant metal in the earth's crust.
3. The properties of non - metals is almost similar with that of metals.
4. Bromine and mercury are non- metallic and metallic liquid elements at room temperature respectively.
5. Silver is the best conductor of heat and electricity.
6. Iodine is solid, bromine is liquid and fluorine and chlorine are gases due to strength of London force.
7. Animals and plants are non - renewable natural resources.
8. Most of the elements are non - metals .
9. The very active metals are manufactured industrially by chemical reduction.
10. The reactivity series is a useful guide to predict the relative reactivity of metals .

II. Matching Items

A	B
11. Al_2O_3	A. Marble
12. Na_3AlF_6	B. Sphalerite
13. Fe_2O_3	C. Galena
14. $FeCO_3$	D. Cinnabar
15. Fe_3O_4	E. Tenorite
16. FeS_2	F. Cuprite
17. $CuFeS_2$	G. Covellite
18. Cu_2S	H. Chalcopyrite
19. CuS	I. Chalcopyrite

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20. Cu_2O J. Fool's gold
21. CuO K. Siderite
22. HgS L. Magnetite
23. PbS M. Hematite
24. ZnS N. Cryolite
25. $CaCO_3$ O. Corundum
P. Slacked lime.

III. Multiple Choice Items

26. Which one of the following is not an ore of iron ?
A. Hematite C. Siderite
B. Magnetite D. Chalcopyrite

27. Which of the following is not the use of nitrogen ?
A. Making matches C. Production of ammonia
B. Making fertilizes D. As a refrigerant

28. Which of the following is correct about the allotropes of phosphorus ?
A. White phosphorus is a waxy and reactive solid
B. White phosphorus is soluble in water
C. White phosphorus is more stable than red phosphorus
D. The allotropes of phosphorus are not inter convertible.

29. Which of the following reaction takes place in the manufacture of aluminum by Hall process ?
A. Aluminum oxides is decomposed by heat
B. Cryolite is electrolyzed at 2000°C
C. Bauxite is electrolyzed using graphite electrode
D. The impurities in the aluminum are removed by limestone.

30. Which of the following is used in the manufacture of phosgene and mustard gas ?
A. N_2O B. CO C. NO_2 D. Cl_2

31. Which of the following elements is stored under water to prevent it from reacting with oxygen ?
A. Magnesium B. Sodium

C. White phosphorus D. Rhombic sulphur

32. Which of the following pairs are not allotropes ?
A. Rhombic and monoclinic sulphur
B. Pig iron and steel
C. Ozone and oxygen
D. White phosphorus and red phosphorus.

33. What is the main commercial source of chlorine ?
A. NaCl B. HCl C. $CaCl_2$ D. KCl

34. Which of the following is the purpose of treating the crude bauxite ore with hot aqueous sodium hydroxide in the extraction of aluminum by Hall process ?
A. To get pure aluminum oxide
B. To decrease the solubility of aluminum oxide
C. To dissolve impurities like iron oxides
D. To lower the melting point of crude bauxite

35. The molten iron from the blast furnace is called ____
A. Wrought iron C. Cast iron
B. Pig iron D. Steel

36. Sterling silver is an alloy of silver and ____
A. Cu B. Zn C. Pb D. Sn

37. In a mixture of an alloy, which property is not affected ?
A. Melting point C. Crystal structure
B. Hardness D. Conductivity

38. The most active non - metallic element is
A. Chlorine C. Oxygen
B. Fluorine D. Sulfur

39. The order of decreasing activity of the halogens is
A. F, Cl, I, Br C. Cl, F, Br, I
B. F, Cl, Br, I D. Cl, Br, I, F

40. Which of the following properties are attributed to most metals ?
I. They are conductors of heat and electricity
II. They are malleable and ductile

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III. They are all solids at room temperature

A. I only
B. II only
C. I and II
D. I, II and III

IV. Short Answer Items

41. List some physical properties of non - metals.

42. _____ are mixtures of two or more metals and are formed by mixing molten metals throughly

43. ____ is a naturally occurring mineral from which a metal can be extracted.

44. The method used to extract a metal from its ore depends on ___

45. List the composition of Pewter

Answer for supplementary exercise on unit 10 part I

1. True	3. False	5. True	7. False	9. False
2. True	4. True	6. True	8. False	10. True

Part II

11. O	13. M	15. L	17. I	19. G	21. E	23. C	25. A
12. N	14. K	16. J	18. H	20. F	22. D	24. B	

Part III

26. D	28. A	30. D	32. B	34. A	36. A	38. B	40. C
27. A	29. C	31. C	33. A	35. B	37. D	39. B	

Part IV

41. They exist in the three states of matter

They are oxidizing agents

They are not malleable and ductile

42. Alloys

43. Ore

44. The position of the metal in the reactivity series.

45. Pewter = 30 % Pb + 70 % Sn + small amount of Sb

UNIT 11

Hydrocarbons and their Natural Source

11.1 Introduction

Carbon is unique. It forms inorganic substances such as carbon dioxide, carbonmonxide, carbonates, hydrogen carbonates, cyanides. It also forms organic substances with out which life could not exist.

Organic chemistry may be defined simply as the chemistry of the compounds of carbon.

In the eary time, organic chemistry deals with substances that were produced by living organisms with the help of **"Vital force"** that could not be manufactured in the laboratory.

☑ **Friedrich Wohler** disproved the idea of vital force theory in 1828 by manufacturing the first syntheitc organic compound **(Urea)** in the laboratory.

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$$NaOCN + NH_4Cl \longrightarrow NH_4OCN + NaCl$$

$$NH_4OCN \longrightarrow \underset{\text{Urea}}{NH_2-\overset{\overset{O}{\|}}{C}-NH_2}$$

Ammonium Cyanate

☑ After Friedrich Wohler, chemists have synthesized thousands of carbon compounds in areas of dyes, plastic, textile fibers, medicines and drugs.

Classification of Organic Compounds

Organic compounds are classified into hydrocarbons and derivative hydrocarbons. **Hydrocarbons** are compounds of hydrogen and carbon only. **Derivative hydrocarbons** are compounds of carbon, hydrogen and other atoms.

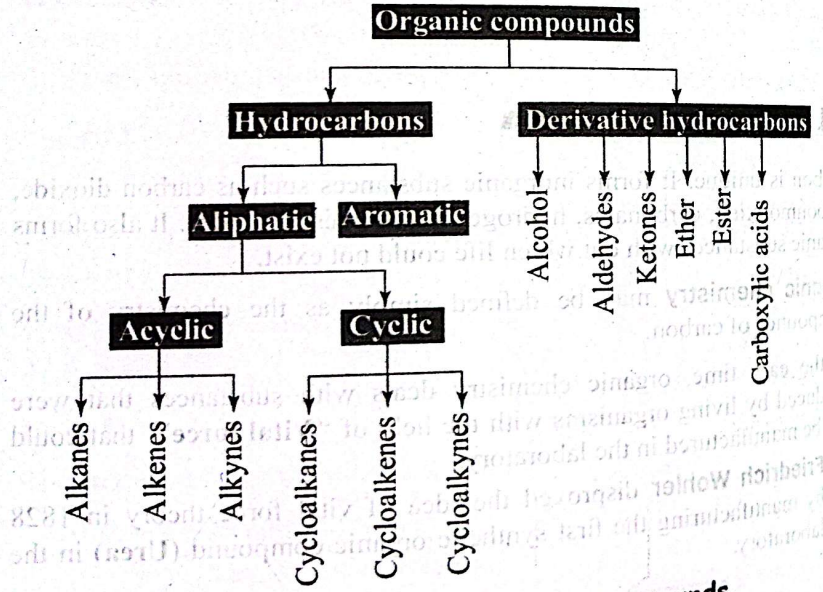


Fig 11. 1 Classification of Organic Compounds

Carbon atom is capable of forming straight chains, branched chains, cyclic compounds and other complicated structure of hydrocarbon. This property of carbon is known as **catenation**.

Functional group is part of the molecule or compound that determines the chemical properties of that molecules or compounds

Different Formula of Organic Compounds

Molecular formula - number atoms of each element expressed with subscripts.

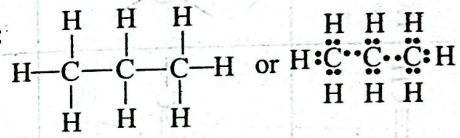
Example

Propane: C_3H_8

Detail structural formula - formula that shows bonding pairs by using line or electron - dot

Example

propane:



Condensed formula - formula that does not show bonds

Example

Propane: $CH_3\ CH_2\ CH_3$

Line formula - formula that does not show carbon and hydrogen atoms.

Prepane: Butane:

Polygon formula - line formula for cyclic hydrocarbons

Example

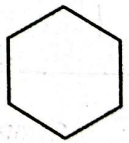
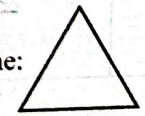
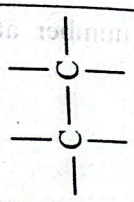
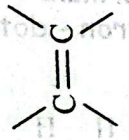
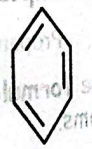
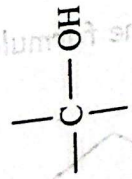
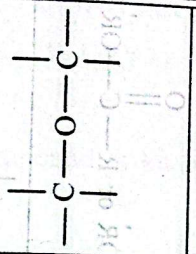
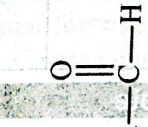
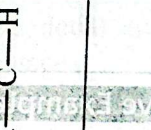
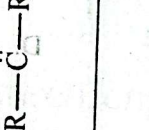
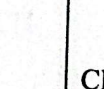
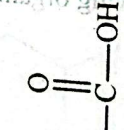
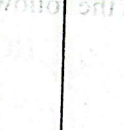
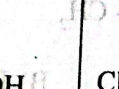
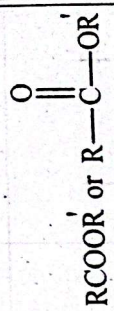
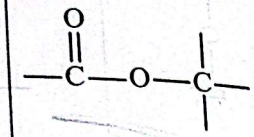
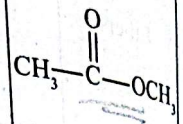
cyclohexane:  cyclopropane: 

Table 11.1 Some Organic Compounds, their Functional Group and Example

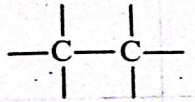
Organic compounds	General formula	Functional group	Example
Alkanes	RH		$CH_3—CH_3$
Alkenes	$RCH=CH_2$		$CH_2=CH_2$
Alkynes	$RC\equiv CH$		$HC\equiv CH$
Aromatics			
Alcohols	R—OH		CH_3CH_2OH

Ethers	R—O—R'		CH_3OCH_3
Aldehydes	 RCHO or 		CH_3 - CHO
Ketones	RCOR' or 		
Carboxylic acids	 RCOOH or 		

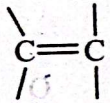
Esters			

Illustrative Example

1. One of the following is the functional group of alkene. Which one is it?

A. 

C.

B. 

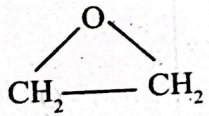
D.

Solution

Alkenes contain carbon - carbon double bond.

Answer: B

2. What functional group present in the following organic compound?



A) Ether B) Ester C) Adehyde D) Ketone

Solution

The presence of C - O - C bonds represent an ether.

Answer: A

3. Which of the following is the first organic compound synthesized from inorganic source?

A) NH_4Cl C) $(NH_2)_2CO$

B) NH_4CNO D) AgCNO

Solution

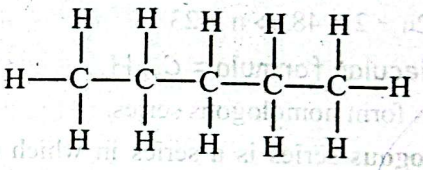
The first synthetic organic compound is known as urea with a formula of $(NH_2)_2 CO$.

Answer: C

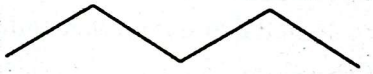
4. Write molecular formula, detial structural formula, condensed formula and line formula for pentane.

Solution

Molecular formula: C_5H_{12}

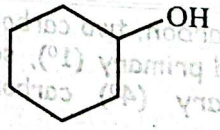
Detail structure: 

Condensed formula: $CH_3\ CH_2\ CH_2\ CH_2\ CH_3$

Line formula: 

5. Write the polygon formula for cyclohexanol.

Solution



11.2 Saturated Hydrocarbons

I. Alkanes

- They are also called **paraffins.**
- They are saturated hydrocarbon (contain carbon - carbon single bond only).
- They have a molecular formula of C_nH_{2n+2} = 1, 2, 3, ...

Example

1. Determine the molecular formula of an alkane with 39 carbon atoms.

Solution

$C = 39 \Rightarrow n = 39$

$C_nH_{2n+2} = C_{39}H_{2\times 39+2} = C_{39}H_{80}$

2. Find the molecular formula of an alkane with48 hydrogen atoms.

Solution

In a give alkane, number of carbon atoms equal to n but number of hydrogen atoms equal to 2n + 2.

Number of H - atoms = 2n + 2

$2n + 2 = 48 \Rightarrow n = 23$

Thus, molecular formula = $C_{23}H_{48}$

- Alkanes form homologous series.
- **Homologous** series is a series in which one member differ from the next by the same repeating cluster of atoms (methylene group - CH_2-)

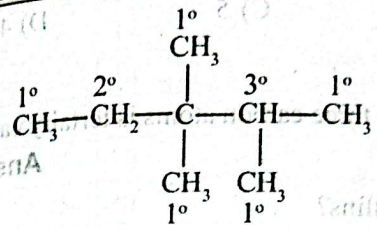
Example

$CH_4, C_2H_6, C_2H_8, C_4H_{10}$, ------

The individual members are called **homologues**

The carbon atoms bounded to one carbon, two carbons, three carbons and four carbons are called primary (1°), secondary (2°), tertiary (3°) and quaternary (4°) carbon atoms respectively.

The hydrogen atoms bonded to 1°, 2° and 3° carbon atoms are 1°, 2° and 3° hydrogen atoms respectively. There is no quaternary hydrogen.



Illustrative Examples

6. An alkane with 15 carbon atoms would have the formula:

A) $C_{15}H_{30}$ B) $C_{15}H_{28}$ C) $C_{15}H_{32}$ D) $C_{15}H_{31}$

Solution

$n = 15 \Rightarrow C_{15}H_{2\times 15+2} = C_{15}H_{32}$ **Answer: C**

7. An alkane that has 26 hydrogen atoms would have a formula of

A) $C_{13}H_{26}$ B) $C_{12}H_{26}$ C) $C_{11}H_{24}$ D) $C_{14}H_{26}$

Solution

$2n + 2 = 26 \Rightarrow n = 12$, Formula = $C_{12}H_{26}$ **Answer: B**

8. Which sets of compounds represent a homologous series?

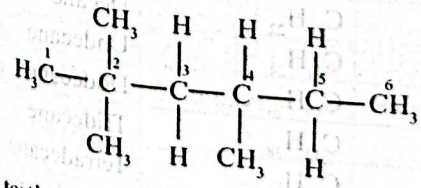
A) C_2H_4, C_3H_8, C_4H_8 C) CH_4, C_2H_6, C_3H_8

B) C_2H_2, C_3H_6, C_4H_8 D) CH_4, C_2H_4, C_3H_6

Solution

In alkanes, the number of hydrogen atom is twice plus two to that of carbon atoms. **Answer: C**

9. Given the structure below



A tertiary carbon atom is shown by the number ____

A) 2 B) 3 C) 5 D) 4

Solution

A carbon atom bonded to three carbon atoms is tertairy carbon .

Answer: D

10. Why alkanes are called paraffins?

Solution

Because they have little affinity towards many reagents. The word paraffin comes from Latin meaning "little affinity"

Nomeclature of Alkanes

The name of alkanes end with the suffix "- **ane**"

Table 11.2 Prefixes, Carbon Number, Molecular Formula and Name

Perfixes	Carbon Number	Molecular Formula	Name
Meth -	1	CH_4	Methane
Eth -	2	C_2H_6	Ethane
Prop -	3	C_3H_8	Propane
But -	4	C_4H_{10}	Butane
Pent -	5	C_5H_{12}	Pentane
Hex -	6	C_6H_{14}	Hexane
Hept -	7	C_7H_{16}	Heptane
Oct -	8	C_8H_{18}	Octane
Non -	9	C_9H_{20}	Nonane
Dec -	10	$C_{10}H_{22}$	Decane
Undec -	11	$C_{11}H_{24}$	Undecane
Dodec -	12	$C_{12}H_{26}$	Dodecane
Tridec -	13	$C_{13}H_{28}$	Tridecane
Tetradec -	14	$C_{14}H_{30}$	Tetradecane
Pentadec -	15	$C_{15}H_{32}$	Pentadecane
Hexade -	16	$C_{16}H_{34}$	Hexadecane

Heptadec -	17	$C_{17}H_{36}$	Heptadecane
Octadec -	18	$C_{18}H_{38}$	Octadecane
Nonadec -	19	$C_{19}H_{40}$	Nonadecane
Eicos -	20	$C_{20}H_{42}$	Eicosane
Triacot -	30	$C_{30}H_{62}$	Triacontane
Tetracont -	40	$C_{40}H_{82}$	Tetracontane
Pentacont -	50	$C_{50}H_{102}$	Pentacontane
Hexacont -	60	$C_{60}H_{122}$	Hexacontane
Heptacont -	70	$C_{70}H_{142}$	Heptacontane
Octacont -	80	$C_{80}H_{162}$	Octacontane
Nonacont -	90	$C_{90}H_{182}$	Nonacontane
Hect -	100	$C_{100}H_{202}$	Hectane

Methane is the simplest alkane, without functional group

Alkyl groups (R-) are radicals obtained from alkane by reducing one hydrogen atom.

The general formula of alkyl group is C_nH_{2n+1}, $n \geq 1$

Example

Determine the molecular formula of an alkyl group with carbon number

a) 1 b) 2 c) 3

Solution

a) $C_nH_{2n+1} = CH_2 \times 1 = CH_3$-

b) $C_nH_{2n+1} = C_2H_{2\times 2+1} = C_2H_5$-

c) $C_nH_{2n+1} = C_2H_{2\times 3+1} = C_3H_7$-

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Table 11.3 Molecular Formula of Alkane, Molecular and Condensed Formula of Alkyl Groups and their Common Names

Molecular formula of alkane	Molecular formula of alkyl group	Condensed formula of alkyl group	Common Name
CH_4	CH_3—	CH_3—	Methyl
C_2H_6	C_2H_5-	CH_3CH_2—	Ethyl
C_3H_8	C_3H_7—	$CH_3CH_2CH_2$—	n - propyl
		$CH_3CH(CH_3)$	Isopropyl
C_4H_{10}	C_4H_9—	$CH_3CH_2CH_2CH_2$—	n - Butyl
		$-CH_2-CH(CH_3)-CH_3$	Iso - butyl
		$CH_3-CH(CH_3)-CH_2-CH_3$	Sec - butyl
		$CH_3-C(CH_3)_2-$	Tert - butyl (t - butyl)

Illustrative Examples

11. The base name of hydrocarbon with 8 carbon atoms is

A) Hept B) Non- C) Octadec- D) Oct-

Answer: D

12. Which one of the following structure represents an iso - butyl alkyl group?

A) $-CH_2-CH(CH_3)-CH_3$ B) $CH_3-CH(CH_3)-CH_2-CH_3$

C) $-C(CH_3)_2-CH_3$ D) $-CH_2-CH_2-CH_2-CH_3$

Answer: A

13. Find the molecular formula of alkyl group with carbon

a) 6 b) 8

Solution

a) $C_nH_{2n+1} = C_6H_{2\times6+1} = C_6H_{13}$ -

b) $C_nH_{2n+1} = C_8H_{2\times8+1} = C_8H_{17}$ -

14. Draw the condensed formula of tert butyl

Solution

$H_3C-C(CH_3)_2-$

A) Common Names of Alkanes

Common name of alkanes take some prefixes to identify one from another.

i. **Normal (n)** - to name unbranched alkanes

Example

$CH_3CH_2CH_2CH_3$ n - Butane

$CH_3CH_2CH_2CH_2CH_3$ n - pentane

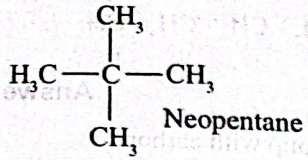
ii. **Iso** - to name alkanes that contain a methyl group on the second carbon

Example

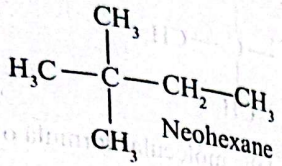
$CH_3-CH(CH_3)-CH_2-CH_3$ Iso pentane

iii. **Neo** - is used for naming alkanes when the central carbon atom is bonded with other four carbon atoms.

Example



Neopentane



Neohexane

B) IUPAC Name of Alkanes

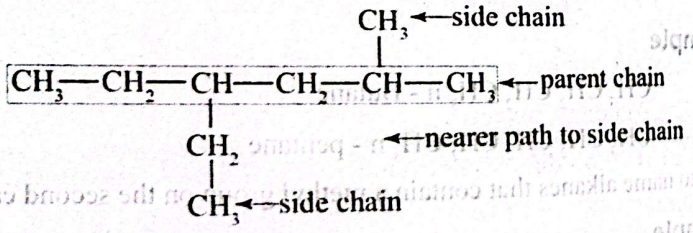
IUPAC - International union pure and Applied chemistry

➥ It is very important to name branched chain hydrocarbons

Rule

1. The name of an alkane ends with the suffix "**-ane**"
2. Identify the longest centinuous chain
3. Count from the end nearer to the side chains
4. Arrange the side chains in alphabetical or size (Molar mass) order.
5. If the same side chain appears twice, three times, four times, use the prefix di, tri, tetra respectively.
6. Number and number are separated by comma (,).
7. Number and letter are separated by hyphen (-).

Example



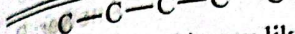
4 - Ethyl - 2 - methyl hexane : Alphabetical order

2 - methyl - 4 - ethyl hexane : size order.

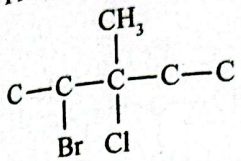
Draw the structural formula for 2 - Bromo - 3 - chloro - 3 - methyl pentane.

Solution

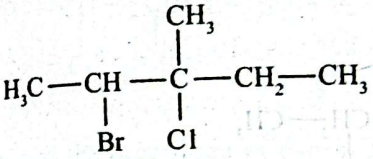
i. Based on the parent name, draw the skeletal structure.



ii. From any side you like, place the side chains on the given position

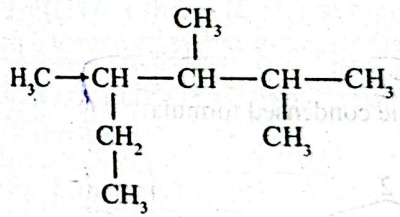


iii. Give the required number of hydrogen atoms for each carbon.



Illustrative Examples

15. The IUPAC name of the compound with the following structure is



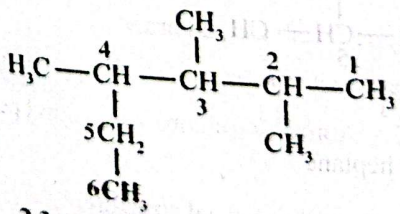
A) 4 - ethyl - 2,3 - dimethyl pentane

B) 2,3,4 - trimethyl hexane

C) 3,4,5 - trimethyl hexane

D) 2 - ethyl - 3,4 - dimethyl pentane

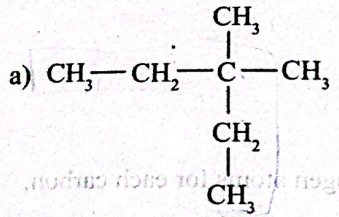
Solution

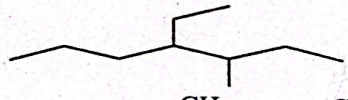


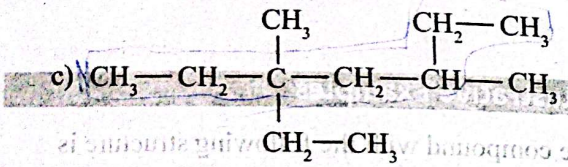
2,3,4 - Trimethyl hexane

Answer: B

16. Give the IUPAC name of:

a) 

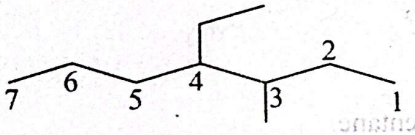
b) 

c) 

Solution

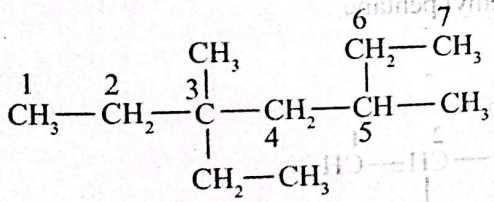
a) 3,3 - dimethyl pentane

b) Follow the same rule as the condensed formula



4 - Ethyl - 3 - methyl heptane

c) If side chains appear on the same carbon number on both sides, give priority for more number of side chains.



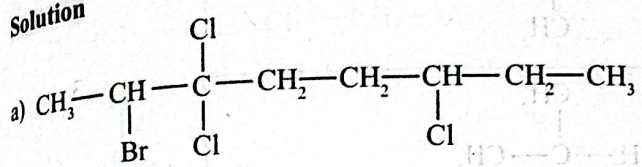
3 - Ethyl - 3,5- dimethyl heptane

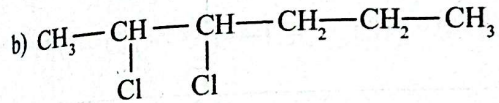
17. Draw the structure of:

a) 2 - Bromo - 3,3,6 - trichloroctane

b) 2,3 - dichloro hexane

Solution

a) 

b) 

Isomerism

Isomerism is the occurrence of two or more compounds with the same molecular formula but different structural formula. The compounds are called **isomers.**

Methane (CH_4), ethane (C_2H_6) and propane (C_3H_8) do not have isomers because they have no different structural formula.

Table 11.4 Isomers of Butane and Pentane with their Structural and Molecular Formula

Alkane	Structure formula	Number of isomers
Butane (C_4H_{10})	$CH_3—CH_2—CH_2—CH_3$	2
	$CH_3—CH(CH_3)—CH_3$	

Pentane (C_5H_{12})	$CH_3—CH_2—CH_2—CH_2—CH_3$	3
	$CH_3—CH(CH_3)—CH_2—CH_3$	
	$CH_3—C(CH_3)_2—CH_3$	

Note: As the number of carbon atoms increases, number of isomers also increases. Alkanes show structural isomers

Illustrative Examples

18. Which type of isomers are common in alkanes?

A) Position isomers C) Geometrical isomers

B) Optical isomers D) Structure isomer

Answer: D

19. Isomers are different in all of the following. **EXCEPT**

A) Chemical property C) Physical property

B) Structure D) Molecular mass

Solution

Isomers have the same molecular formula and molecular mass but different structure, physical and chemical property.

Answer: D

20. How many different structure arrangments are possible for heptane?

A) 3 B) 5 C) 9 D) 7

Solution

Number of isomers for simple alkanes $= 2^{n-4} + 1$

n = number of carbon atoms.

$= 2^{n-4} + 1 = 2^{7-4} + 1 = 9$

Answer: C

21. Draw the five isomers of hexane (C_6H_{14}) and write the IUPAC name for each

Solution

$CH_3—CH_2—CH_2—CH_2—CH_2—CH_3$ Hexane

$CH_3—CH(CH_3)—CH_2—CH_2—CH_3$ 2 - methyl pentane

$CH_3—CH_2—CH(CH_3)—CH_2—CH_3$ 3 - methyl pentane

$CH_3—CH(CH_3)—CH(CH_3)—CH_3$ 2, 3 - dimethyl butane

$CH_3—C(CH_3)_2—CH_2—CH_3$ 2, 2 - dimethyl butane

Physical Properties of alkanes

- ☑ Alkanes are colorless hydrocarbons
- ☑ The first four $(C_1 - C_4)$ alkanes are gases
- ☑ From pentane (C_5H_{12}) to heptadecane $(C_{17}H_{36})$ are liquids at room temperature.
- ☑ Octadecane $(C_{18}H_{38})$ and above exist in solid state at room temperature
- ☑ As carbon number increases, boiling point also increases.

Example

Boiling point: $CH_4 < C_2H_6 < C_3H_8 < .....$

- ☑ As branching or side chains increases, boiling point decreases, because of the molecules approaches that of sphere which results in low intermolecular force of attraction.

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Example

Boiling of point: $CH_3—CH_2—CH_2—CH_2—CH_3$

$> CH_3—\underset{\displaystyle CH_3}{\overset{\displaystyle |}{CH}}—CH_2—CH_3 > H_3C—\underset{\displaystyle CH_3}{\overset{\displaystyle CH_3}{\overset{|}{\underset{|}{C}}}}—CH_3$

☑ Alkanes are insoluble in water because they are non - polar covalent molecules.

☑ Alkanes are generally volatile. However, as carbon number increases, volatility decreases.

Illustrative Examples

22. Which of the following compounds would be in the liquid state at room temperature?

A) CH_4 B) C_6H_{14} C) C_3H_8 D) $C_{25}H_{52}$

Solution

C_5 - C_{17} exist in liquid state

Answer: B

23. Which of the following is the physical property of alkanes?

A) Their densities decrease with increasing molecular mass along the homologous series.

B) Their boiling points decrease with carbon number

C) The first four alkanes are solid at room temperature

D) They are colorless

Answer: D

24. Alkanes are insoluble in water because

A) They are solid substances

B) They have strong intermolecular affraction forces.

C) They are non - polar substances.

D) They possess carbon - carbon single bond.

Solution

Due to the "like dissolves like" principle alkanes are non - polar and are insoluble in polar solvent, water.

Answer: C

25. Which of the following physical properties decrease as the number of carbon atom increases in a homologous series of alkanes?

A) Intermolecular force C) Melting point

B) Density D) Solubility

Answer: D

26. Arrange the following alkanes in increasing order of boiling point.

I. $CH_3—CH_2—CH_2—CH_2—CH_2—CH_3$

II. $CH_3—\underset{\displaystyle CH_3}{\overset{\displaystyle |}{CH}}—CH_2—CH_2—CH_3$

III. $CH_3—\underset{\displaystyle CH_3}{\overset{\displaystyle CH_3}{\overset{|}{\underset{|}{C}}}}—CH_2—CH_3$

Solution

As branching increases, boiling points decrease. III < II < I

Preparation of Alkanes

☑ The natural sources of alkanes are **pertoleum and natural gas.**

☑ The difference in boiling points of mixtures of the liquid alkanes found in petroeum make it possible to separate the various components by **fractional distillation.** This is the major industrial process used in refining petroleum into gasoline, kerosene, lubricating oils and several other minor components.

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In the laboratory alkanes can be prepared by

1. Hydrogenation of alkenes

Hydrogenation is addition of hydrogen to a substance.

$$\text{Alkenes} + \text{hydrogen} \xrightarrow[heat]{Catalyst} \text{Alkanes}$$

$$\underset{\text{Alkene}}{R\text{-}CH=CH_2} + H_2 \xrightarrow[heat]{Pt} \underset{\text{Alkane}}{R\,CH_2\text{-}CH_3}$$

Example

$$CH_2=CH_2 + H_2 \xrightarrow[heat]{Pt} CH_3\text{-}CH_3$$

Note: This method is not used to prepare methane

2. Wurtz synthesis

It is the reaction between haloalkane and sodium to yield alkane and sodium halide.

$$\underset{\substack{\text{Alkylhalide}\\(\text{haloalkane})}}{2R\text{—}X} + 2Na \longrightarrow \underset{\text{Alkane}}{R\text{—}R} + 2NaX \quad , X \text{ halogen}$$

Example

$$2CH_3I + 2Na \longrightarrow CH_3\text{—}CH_3 + 2NaI$$

3. Heating sodium salt of organic acid with soda lime.

Soda lime is a mixture sodium hydroxide and calcium oxide.

Example

$$\underset{\substack{\text{Sodium}\\\text{carboxylate}}}{RCOONa} + NaOH \xrightarrow[heat]{CaO} \underset{\text{Alkane}}{RH} + Na_2CO_3$$

Example

$$CH_3CH_2COONa + NaOH \xrightarrow[heat]{CaO} CH_3\text{—}CH_3 + Na_2CO_3 + H_2O$$

Sodium propanoate

Note:- This method is used to prepare methane

- It can be also prepared from fractional distillation of petroleum.
- It is formed by the decay of animals and plants in marshy lands.

Illustrative Examples

27. All of the following reactions produce an alkane. **EXCEPT.**

A) $H_3COONa + NaOH \longrightarrow$

B) $CH_2=CH_2 + H_2 \xrightarrow[catarlyst]{heat}$

C) $2CH_3I + 2Na \rightarrow$

D) $CH_3 - CH = CH + H_2O \xrightarrow{H_2SO_4}$

Solution

$$H_3COONa + NaOH \rightarrow \underset{\text{Methane}}{CH_4} + Na_2CO_3$$

$$CH_2=CH_2 + H_2 \xrightarrow[catarlyst]{heat} CH_3\text{-}CH_3 \text{ (Ethane)}$$

$$2CH_3I + 2Na \rightarrow CH_3\text{-}CH_3 + 2NaI$$

$$CH_3\text{—}CH=CH_2 + H_2O \xrightarrow{H_2SO_4} \underset{\substack{\text{Propanol(alcohol)}\\ OH}}{CH_3\text{—}CH\text{—}CH_3}$$

Answer: D

28. Which of the following reactions could produce alkanes in the laboratory?

A) $R\text{—}CH=CH_2 + H_2 \xrightarrow[heat]{Pt}$

B) $R\text{—}CH_2\text{—}CH_2\text{—}OH + H_2SO_4 \xrightarrow{heat}$

C) $R\text{—}CH_2\text{—}CH_2\text{—}X + KOH \longrightarrow$

D) $R\text{—}CH=CH_2 + H_2O \longrightarrow$

Answer: A

29. Industrially alkanes are produced by _____

Answer: Fractional distillation of petroleum

30. **What are raw materials used by Wurtz synthesis?**

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Solution

Halogenated alkane and sodium metal.

Chemical Preparties of Alkanes

Alkanes have little affinity towards chemical reagents. However, they undergo the following chemical reactions.

1. Substitution Reaction is the Typical Reaction of Alkanes

- **Halogenations** of alkanes is a substitution reaction.
- The reaction of bromine and chlorine with alkanes occur in the presence of sunlight, such type of reaction is called **photochemical reaction.**
- **Bromination** of methane may produce different products depending on the amount of bromine.

$$CH_4 + Br_2 \xrightarrow{light} CH_3Br + HBr$$

$$CH_3Br + Br_2 \rightarrow CH_2Br_2 + HBr$$

$$CH_2Br + Br_2 \rightarrow CHBr_3 + HBr$$

$$CHBr_3 + Br_2 \rightarrow CBr_4 + HBr$$

over all reaction

$$CH_4 + 4Br_2 \rightarrow CBr_4 + HBr$$

The sequence of steps involved in the reaction is called **reaction mechanism.**

For bromination of methane, the steps involved are:

i. Chain Initiating Step

It involves production of reactive species called **free radicals.**

$$Br{-}Br \xrightarrow{light} 2Br\cdot$$

ii. Chain Propagative Step

This step consumes and generates free radicals

$$Br\cdot + CH_4 \longrightarrow CH_3^{\cdot} + HBr$$

iii. Chain Terminating Step

Here free radicals are consumed but not generated.

$$CH_3^{\cdot} + Br\cdot \longrightarrow CH_3Br$$

2. Elimation Reaction

☑ Alkanes undergo elimation reaction at relatively high temperature.

☑ It involves removal of small molecules.

Example

$$CH_3{-}CH_2{-}CH_3 \xrightarrow{heat} CH_2{=}CH_2 + CH_4$$

3. Combustion Reactions

Alkanes + excess air → Carbondioxide + water + heat

Alkanes + limited air → carbonmonoxide + water +heat

Example

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + heat$$

$$2CH_4 + 3O_2 \rightarrow 2CO + 4H_2O + heat$$

Uses of Alkanes

☑ Used as **fuel** for heating, cooking, electricity generation and transportation.

☑ Used as solvents for non - polar substance, lubricants for machinery and engines.

☑ Waxes and paraffins for candies, cosmetics, and food packaging.

☑ Asphalt for road construction.

Illustrative Examples

31. Which of the following is an elimation reaction?

A) $CH_3CH_2CH_3 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

B) $CH_3CH{=}CH_2 + H_2 \xrightarrow[heat]{Pt} CH_3CH_2CH_3$

C) $CH_3CH_2CH_3 \xrightarrow{heat} CH_3CH{=}CH_2 + H_2$

D) $CH_3CH_2CH_3 + Cl_2 \xrightarrow{UV} CH_3CH_2CH_2Cl + HCl$

Answer: C

32. The process of converting higher alkanes into smaller alkanes and

alkenes to increase the gasoline yield is known as

A) Alkylation C) Cracking

B) Polymerization D) Saponification

Solution

Cracking is a special type of elimination reaction that converts large alkanes into smaller alkenes and alkanes.

Answer: C

33. The possible reason why alkanes are relatively stable to most common laboratory reagents at room temperature is due to

A) High strength of carbon to carbon and hydrogen to carbon bonds.

B) Their strong intermolecular forces

C) Their insolubility in water

D) Their molecular masses.

Solution

Alkanes are non - polar saturated hydrocarbons and inert towards many laboratory reagents such as acids, bases, oxidizing agents and reducing agents.

Answer: C

34. Which of the following is the structure for methyl chloride?

A) CH_3Cl B) CH_2Cl_2 C) $CHCl_3$ D) CCl_4

Solution

CH_4 undergoes chlorination reaction

$$CH_4 + Cl_2 \xrightarrow{heat} CH_3Cl + HCl$$

Methyl chloride or chlorom methane = CH_3Cl

$$CH_3Cl + Cl_2 \rightarrow CH_2Cl_2 + HCl$$

Methylene chloride or dichlorom ethane = CH_2Cl

$$CH_2Cl_2 + Cl_2 \rightarrow CHCl_3 + HCl$$

Chloro form or Trichloro methane = $CHCl_3$

$$CHCl_3 + Cl_2 \rightarrow CCl_4 + HCl$$

Carbon tetrachloride or tetrachloro methane = CCl_4

Answer: A

35. Alkanes are primary used as

A) Fuels C) Raw materials to male soaps

B) Organic D) Raw materials to make plastics

Answer: A

36. List the uses of alkans as raw materials

Solution

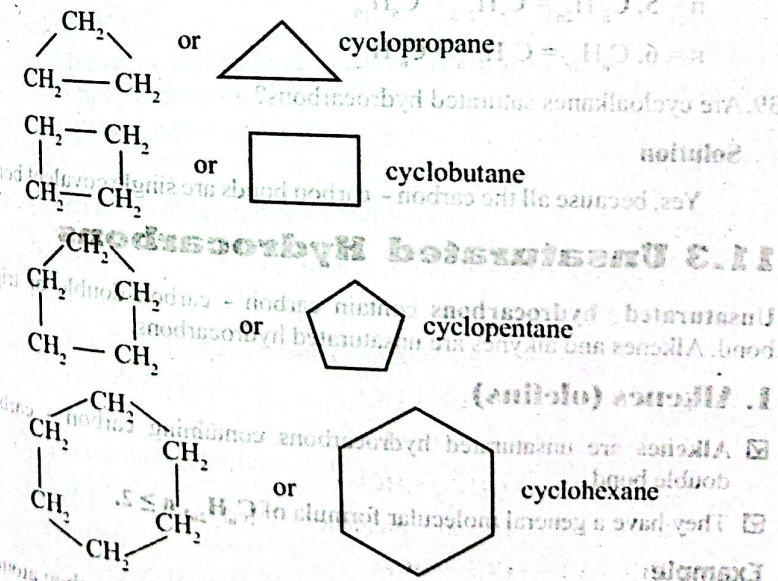
Alkanes are used as raw materials in plastics, soaps, alcohols and other chemical industries.

2. Cycloalkanes

☑ They are monocyclic saturated hydrocarbons. The carbon atoms of the molecules are arranged in the form of a ring.

☑ Their molecular formula is represented by C_nH_{2n}, $n \geq 3$.

Example



Note: Cycloalkanes and open chain alkenes are isomers to each other with same carbon atoms.

Illustrative Examples

37. Which of the following is the molecular formula of a cycloalkane?

A) C_4H_{10} B) C_4H_9 C) C_4H_6 D) C_4H_8

Solution

The molecular formula of a cycloalkane should fulfill the general formula C_nH_{2n}, $n \geq 3$.

Answer: D

38. Write the molecular formula of the first four cycloalkanes in the homologous series.

Solution

$n = 3, C_nH_{2n} = C_3H_{2\times3} = C_3H_6$

$n = 4, C_nH_{2n} = C_4H_{2\times4} = C_4H_8$

$n = 5, C_nH_{2n} = C_5H_{2\times5} = C_5H_{10}$

$n = 6, C_nH_{2n} = C_6H_{2\times6} = C_6H_{12}$

39. Are cycloalkanes saturated hydrocarbons?

Solution

Yes, because all the carbon - carbon bonds are single covalent bonds

11.3 Unsaturated Hydrocarbons

Unsaturated hydrocarbons contain carbon - carbon double or triple bond. Alkenes and alkynes are unsaturated hydrocarbons.

1. Alkenes (olefins)

☑ **Alkenes** are unsaturated hydrocarbons containing carbon - carbon double bond.

☑ They have a general molecular formula of C_nH_{2n}, $n \geq 2$.

Example

1) Determine the molecular formula of an alkene with 5 carbon atoms.

Solution

Carbon atoms 5 means $n = 5$

$\Rightarrow C_nH_{2n} = C_{5\ H5\times2} = C_5H_{10}$

2) Find the molecular formula of an alkene that contains 28 hydrogen atoms.

Solution

Number of H - atoms $= 2n = 28 \Rightarrow n = 14$

Thus, the molecular formula of the alkene is $C_{14}H_{28}$

☑ **Alkenes** form a homologous series: C_2H_4, C_3H_6, C_4H_8 -------

☑ They have two fewer hydrogen atoms than the corresponding alkanes.

Illustrative Examples

40. Which of the following general formula represents the group of olefines? (ESSLCE)

A) $C_nH_{2n} + 2, n \geq 1$ C) $C_nH_{2n}, n > 2$

B) $C_nH_{2n}, n \geq 2$ D) $C_nH_{2n} + 1, n \geq 1$

Solution

Alkenes or olefins have a general formula of C_nH_{2n}, $n \geq 2$ because to have carbon - carbon a double bond there must be at least two carbon atoms.

Answer: B

41. Which of the following is the third alkene?

A) C_2H_4 B) C_3H_6 C) C_4H_8 D) C_5H_{10}

Solution

The first alkene contains two carbon atoms. Thus, the third has four carbon atoms.

Answer: C

42. Find the molecular formula of an alkene with carbon number

a) 10 b) 22

Solution

a) $C_nH_{2n} = C_{10}H_{2\times10} = C_{10}H_{20}$

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b) $C_nH_{2n} = C_{22}H_{2\times22} = C_{22}H_{44}$

43. Determine the molecular formula of an alkene that contains 38 hydrogen atoms.

Solution

Number H - atoms = 2n ⇒ n = 19. Thus , the molecular formula of the alkene is $C_{19}H_{38}$.

Nomenclature of Alkenes

A. Common Name of Alkenes

To give common name of an alkene, simply add the suffix -"ylene" to the base name the hydrocarbon.

Common name = base name + ylene

Example

C_2H_4 - Ethylene

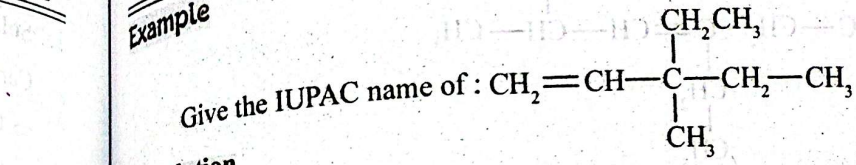
Table 11.5 Some Alkenes and their Common Names

Alkene	Common name
C_2H_4	Ethylene
C_3H_6	Propylene
C_4H_8	Butylene
C_5H_{10}	Pentylene
C_6H_{12}	Hexylene

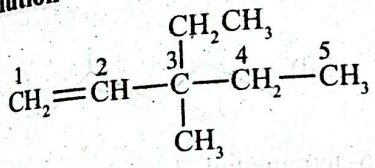
B. IUPAC Name of Alkenes

- ☑ The IUPAC names of alkenes end with the suffix "**-ene**'.
- ☑ Start from the end **nearer to** the double bond and state the position of the double bond.
- ☑ The rest of the rules are similar with that of alkanes.

Example

Give the IUPAC name of : 

Solution

 longest chain (parent structure)

3 - Ethyl - 3 - methyl - 1 - pentene or

3 - Ethyl - 3 - methyl pent - 1 - ene

Illustrative Examples

44. The common name of the second alkene is

A) Butylene B) Ethylene C) Pentylene D) Propylene

Solution

The second alkene has 3 carbon atoms. Therefore, its common name is propylene.

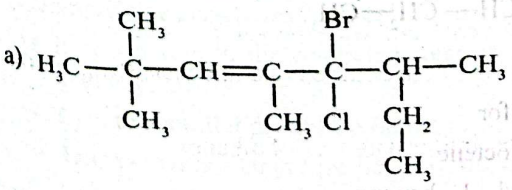
Answer: D

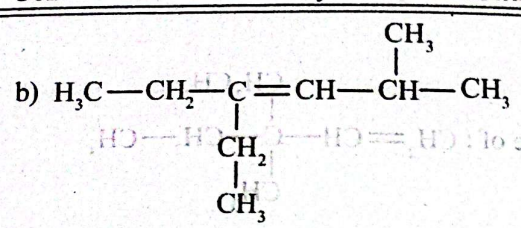
45. Which suffix is added to the base name of a hydrocarbon to give IUPAC name of alkenes?

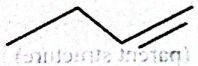
A) -ane B) -ylene C) -ene D) -yne

Answer: C

46. Give the IUPAC name of the following alkenes.

a) 

b) 

c) 

d)

Solution

a) 5 - Bromo - 5 - chloro - 2,2,4,6- tetramethyl -3-octene

(5 - Bromo - 5 - chloro- 2,2,4,6- tetramethyl oct - 3 - ene)

b) Here, the double bond appears in the same carbon number on both sides, start from end nearer to the side chain.

⇒ 4 - Ethyl - 2 - methyl - 3 - hexene

(4 - ethyl - 2 - methyl hex - 3 -ene)

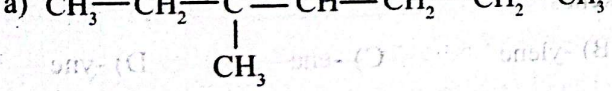
c) 1 - Butene or but - 1 - ene

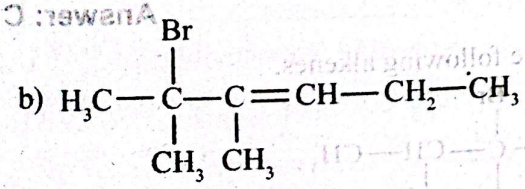
d) 2 - methyl - 2 - pentene or 2 - methylpent - 2 - ene

47. Draw the structural formula of

a) 3 - methyl - 3 - heptene b) 2-Bromo-2,3-dimehylhex-3-ene

Solution

a) 

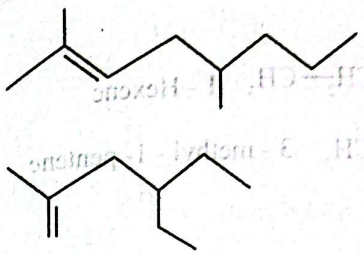
b) 

48. Draw the line formula for

a) 2,5 - dimethyl - 2 - octene

b) 4 - ethyl - 2 - methyl - 1 - hexene

Solution

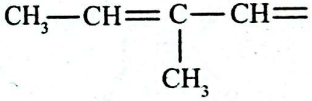


Dienes are unsaturated hydrocarbons that contain two double bonds. Their general formula is C_nH_{2n-2}, $n \geq 3$

The name of dienes = Alka + diene and state the position of the double bonds.

Example

$CH_2=CH-CH=CH_2$ 1, 3 - butadiene

 3 - methyl - 1, 3 - pentadiene

Isomerism in Alkenes

Ethene and **propene** do not have isomers because they have one possible structure.

Alkenes exhibit three isomerisms

1. **Structural isomerism** - change in parent name
2. **Position isomerism** - change in position of double bond
3. **Geometric or cis - trans isomerism** - is a type of stereoismerism.
   - It occurs due to the restricted rotation about the carbon - carbon double bonds.
   - **Cis** - two similar groups lie on the same side of double bond.
   - **Trans** - two similar groups lie on the opposite side of the double bond

Example

1. Draw the structure and position isomers of hexene (C_6H_{12})

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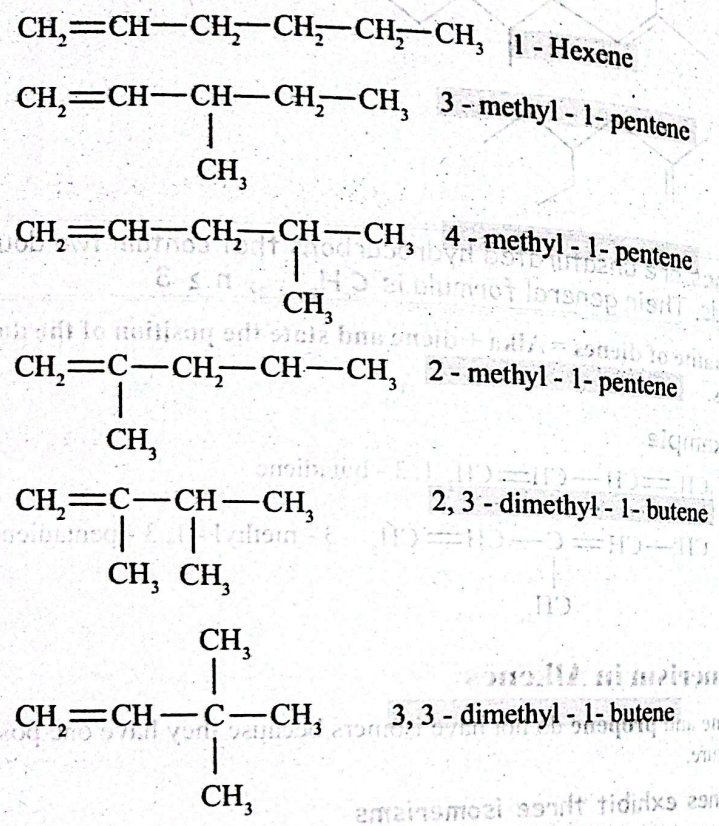
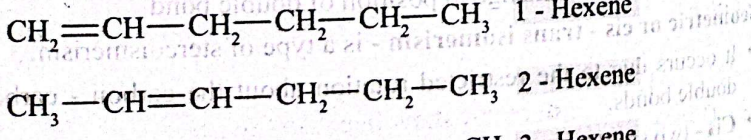
Solution

Structural Isomers

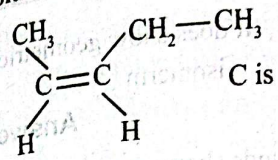

CH2=CH—CH2—CH2—CH2—CH3 1 - Hexene
CH2=CH—CH(CH3)—CH2—CH3 3 - methyl - 1- pentene
CH2=CH—CH2—CH(CH3)—CH3 4 - methyl - 1- pentene
CH2=C(CH3)—CH2—CH—CH3 2 - methyl - 1- pentene
CH2=C(CH3)—CH(CH3)—CH3 2, 3 - dimethyl - 1- butene
CH2=CH—C(CH3)2—CH3 3, 3 - dimethyl - 1- butene

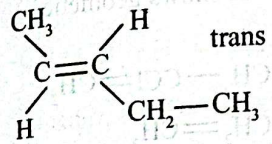
Position isomers

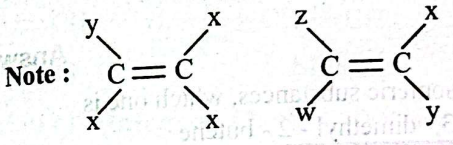


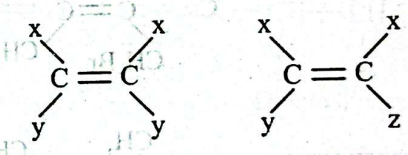
2. Draw the geometric isomers of 2 - pentene.

Solution







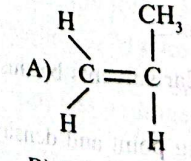


Illustrative Example

49. Which of the following organic compounds shows geometric (cis - trans) isomerism?

A) 1 - propene B) 2 - pentyne C) 3 - hexene D) 1 - butene

Solution



B) It is not an alkene

C) $CH_3—CH_2—\underset{H}{\overset{H}{C}}=C—CH_2—CH_3$ It does show geometric isomerm

Answer: C

50. Which of the following organic compounds shows geometic isomerism?

A. $CH_3—CH=CH—CH_3$
B. $CH_3—CH=CH_2$
C. $CH_3—CCl=CH_2$
D. $CH_2=CH_2$

Answer: A

51. Among the following isomeric substances, which one is Cis - 1,4 - dibromo - 2,3 - dimethyl - 2 - butene

A. CH_2Br, CH_3, $C=C$, CH_2Br, CH_3

C. CH_3, CH_2Br, $C=C$, CH_2Br, CH_3

B. CH_3, CH_3, $C=C$, CH_2Br, CH_2Br

D. CH_3, CH_2Br, $C=C$, CH_2CH_3, Br

Answer: B

Physical Properties of Alkenes

- ☑ C_2 - C_4 are gases
- ☑ C_5 - C_{17} occur in liquid state
- ☑ Greater than or equal to C_{18} exist in solid state.
- ☑ They are insoluble in water but soluble in non polar solvents because they are non - polar.
- ☑ As carbon number increases, boiling point, melting point and density increases.

Chemical properties of Alkenes

1. **Addition reaction** is the most common reaction of alkenes because they are unsaturated hydrocarbons.

a) Addition of Hydrogen (hydrogenation)

Example

$$R\text{-}CH=CH_2 + H_2 \xrightarrow{Pt} RCH_2CH_3$$

Alnene — heal Alkane

$$CH_2=CH_2 + H_2 \xrightarrow[heat]{Pt} CH_3—CH_3$$

Ethene — **Ethane**

b) Addition of Halogens (Halogenation)

➥ Addition of halogens to alkenes yields **vicinal dihaloakans.**

$$R—CH=CH_2 + X_2 \longrightarrow R—\underset{X}{CH}—\underset{X}{CH_2} \quad (X = Cl, Br)$$

Alkene — Dihaloalkane

Example

$$CH_2=CH_2 + Cl_2 \longrightarrow ClCH_2—CH_2Cl$$

Ethene — **1, 2 - di chloroethane**

Note: Bromin water (Br_2 in CCl_4) is used to test unsaturation, When it is added to alkenes, the redish brow color disappears.

c) Addition of hydrogen halide (Hydro halogenation)

☑ This reaction follows the **Marko vinkov's rule.**

☑ According to **Marko Vinkov's rule**, the hydrogen move to the carbon that contains more hydrogen and the halogen toward the carbon that contains less hydrogen. In short, it means "Those that have get more".

$$R—CH=CH_2 + H—X \longrightarrow R—\underset{X}{CH}—CH_3$$

Alkene — **Halo alkane**

Example

$CH_3-CH=CH_2 + HCl \longrightarrow CH_3-\underset{Cl}{CH}-CH_3$

Propene — 2 - chloropropane

d) Addition of water (Hydroation)

This reaction also obeys **Marko vinkov's rule**

$R-CH=CH_2 + H-OH \xrightarrow{H^+} R-\underset{OH}{CH}-CH_3$

Alkene — Alcohol

Example

$CH_3-CH=CH_2 + H_2O \xrightarrow{H^+} CH_3-\underset{OH}{CH}-CH_3$

Propene — 2 - propanol

Note: The most important acid used as catalyst in organic chemical reactions is sulphuricacid.

e) Polymerization (Self - addition) is formation of large molecules (Polymers) from smaller molecules (monomers).

$nCH_2=CH_2 \xrightarrow[\text{catalyst}]{\text{heat}} -(CH_2-CH_2)_n-$

Ethylene (monomer) — **poly ethylene (polymer)**

f) Oxidation of alkenes with strong oxidizing agent

- This reaction yields **diols.**
- It is also used to test unsaturation.

$R-CH=CH_2 \xrightarrow[KMnO_4]{\text{alkaline}} R-\underset{OH}{CH}-\underset{OH}{CH_2}$

Example

$CH_3-CH=CH_2 \xrightarrow[KMnO_4]{\text{alkaline}} CH_3-\underset{OH}{CH}-\underset{OH}{CH_2}$

Illustrative Examples

52. What would be the product X in the chemical reaction shown below?

$CH_3-CH=CH_2 + HBr \rightarrow X$

A. $CH_3-CBr_2-CH_3$

B. $CH_3-\underset{Br}{CH}-CH_2Br$

C. $CH_3-CH_2-CH_2-Br$

D. $CH_3-\underset{Br}{CH}-CH_3$

Solution

$CH_3-CH=CH_2 + H-Br \longrightarrow CH_3-\underset{Br}{CH}-CH_3$

Answer: D

53. Which of the following chemical reactions is a confirmatory test for unsaturation?

A. $CH_2=CH_2 + 3O_2 \longrightarrow 2CO_2 + 2H_2O$

B. $CH_2=CH_2 + Br_2 \longrightarrow CH_2Br-CH_2Br$

C. $CH_2=CH_2 + H_2 \longrightarrow CH_3-CH_3$

D. $nCH_2=CH_2 \longrightarrow -(CH_2-CH_2)_n-$

Solution

Bromination is used to test unsaturation (presence of double or triple bond)

Answer: B

54. $R-CH=CH_2 + X \xrightarrow{Y} R-\underset{OH}{CH}-CH_3$

Which of the following can correctly represent X and Y respectively.

A) H_2O and H_2SO_4

B) O_2 and Ni

C) Ag^+ and H_2O D) OH^- and yeast

Answer: A

55. Which of the following compounds can undergo bromination reaction readily?

A) C_3H_{18} B) C_4H_8 C) C_5H_{12} D) C_6H_{14}

Solution

Bromination is an addition reaction which corresponds to alkene.

Answer: B

56. If $CH_3 - CH = CH_2$ is treated with HCl, the name of the product will be:

A) Propane C) 1 - poholoro propene

B) 1 - chloropropane D) 2 - propane

Solution

$$CH_3—CH=CH_2 + HCl \longrightarrow CH_3—\underset{Cl}{CH}—CH_2$$

2 - chloro propane

Answer: D

57. Complete and balance the following

a. $CH_3—CH_2—\underset{CH_3}{C}=CH_2 + HBr \longrightarrow$

b. $CH_3—\underset{CH_3}{C}=\underset{CH_3}{C}—CH_3 + Br_2 \longrightarrow$

c. $CH_3—\underset{CH_3}{C}=CH_2 + H_2O \longrightarrow$

d. $CH_3—CH=CH—CH_3 \xrightarrow[KMnO_4]{Alkaline}$

Solution

a. $CH_3—CH_2—\underset{CH_3}{C}=CH_2 + HBr \longrightarrow CH_3—CH_2—\underset{CH_3}{\overset{Br}{C}}—CH_3$

2- bromo - 2 - methyl butane

b. $CH_3—\underset{CH_3}{C}=\underset{CH_3}{C}—CH_3 + Br_2 \longrightarrow CH_3—\underset{CH_3}{\overset{Br}{C}}—\underset{CH_3}{\overset{Br}{C}}—CH_3$

2, 2 - dibromo -2, 3 - dimethyl butane

c. $CH_3—\underset{CH_3}{C}=CH_2 + H_2O \longrightarrow CH_3—\underset{CH_3}{\overset{OH}{C}}—CH_3$

2 - methyl -2 - propanol

d. $CH_3—CH=CH—CH_3 \xrightarrow[KMnO_4]{Alkaline} CH_3—\underset{OH}{CH}—\underset{OH}{CH}—CH_3$

2. Combution Reaction

Alkene + oxygen (excess) → carbondioxide + water + heat

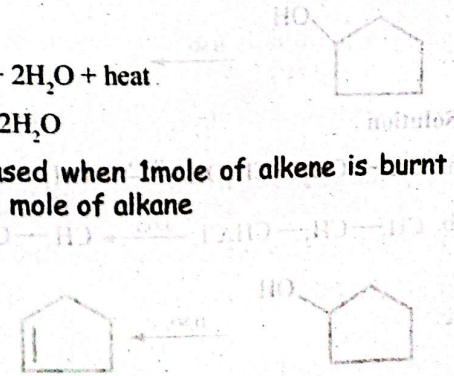
Alkene + oxygen (limited) → Carbonmonoxide + water + heat

Example

$C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$ + heat

$C_2H_4 + 2O_2 \rightarrow 2CO + 2H_2O$

Note: The heat released when 1mole of alkene is burnt is greater than that of 1 mole of alkane



Preparation of Alkenes

1. **Dehydration of alcohols** - is the removal of water a compound by using dehydration agent (H_2SO_4)

$$R{-}CH_2{-}CH_2OH \xrightarrow[heat]{H_2SO_4} R{-}CH{=}CH_2 + H_2O$$

Alcohol **Alkene**

Example

$$CH_3{-}CH_2OH \xrightarrow[heat]{H_2SO_4} CH_2{=}CH_2 + H_2O$$

Ethanol **Ethene**

2. **Dehydro alogenation of alkyl halide with base** the removal of hydrogen halide from a compound.

$$R{-}CH_2{-}CH_2{-}X + KOH \longrightarrow R{-}CH{=}CH_2 + KX + H_2O$$

Alkyl halide **Alkene**

Example

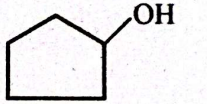
$$CH_3{-}CH_2Cl + KOH \longrightarrow CH_2{=}CH_2 + KCl + H_2O$$

Illustrative Examples

58. Complete and balance the following

a. $CH_3{-}CH_2{-}CH_2OH + H_2SO_4 \xrightarrow{heat}$

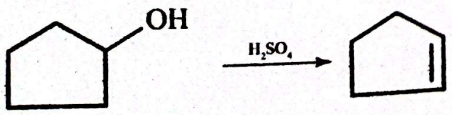
b. $CH_3{-}CH_2{-}CH_2Cl + KOH \longrightarrow$

c.  $\xrightarrow{H_2SO_4}$

Solution

a. $CH_3{-}CH_2{-}CH_2OH \xrightarrow{H_2SO_4} CH_3{-}CH{=}CH_2 + H_2O$

b. $CH_3{-}CH_2{-}CH_2Cl \xrightarrow{KOH} CH_3{-}CH{=}CH_2 + KCl + H_2O$

c. 

59. Write the structural formula of alkyl halide used to make ethene.

Solution

$$CH_3{-}CH_2X + KOH \longrightarrow CH_2{=}CH_2 + KX + H_2O$$

Haloethane (X = Cl , or Br)

Uses of Alkenes

Alkenes are mainly used to make polymers.

Example

$$nCH_2{=}CH_2 \xrightarrow[heat]{catalyst} \left(CH_2{-}CH_2\right)_n$$

poly ethylene

$$nCH_2{=}CH{-}CH_3 \xrightarrow[heat]{catalyst} \left(CH_2{-}\underset{\displaystyle CH_3}{\underset{|}{CH}}\right)_n$$

propylene **polypropylene**

$$nCH_2{=}CHCl \xrightarrow[heat]{catalyst} \left(CH_2{-}\underset{\displaystyle Cl}{\underset{|}{CH}}\right)_n$$

vinyl chloride **polyvinylchloride (PVC)**

Note: $CH_2{=}CH{-}$ vinyl

$CH_3{-}CH_2{=}CH{-}$ allyl

- ☑ **Ethene** is also used for food preservation and artificial ripening of fruits.
- ☑ Alkenes undengo acid - cataryzed hydration to form alcohols.
- ☑ They are used to prepare alkyl halide.

Illustrative Examples

60. Which of the following is the momomer of polyvinyl chloride (PVC)?

A) C_2H_2 B) C_3H_5 C) C_2H_3Cl D) C_2H_5Cl

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Solution

Chloroethene or vinyl chloride (C_2H_3Cl) is the monomer of PVC.

Answer: C

61. The most important industrial application of alkenes is

A) As fuel
B) As solvent
C) In polymer production
D) In insecticide production

Answer: C

62. List the uses of ethene as raw material

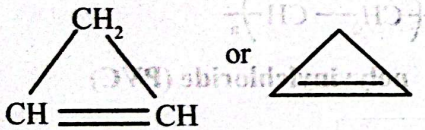
Solution

To make ethanol, ethylene oxide, ethanal and polyethylene.

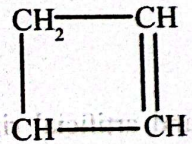
Cyclo Alkenes

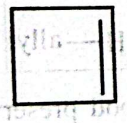
- They are cyclic alkenes
- They are unsaturated hydrocarbons
- Their general formula is C_nH_{2n-2}, $n \geq 3$

Example

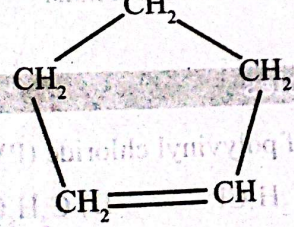


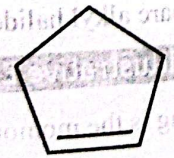
or or C_3H_4 cyclopropene



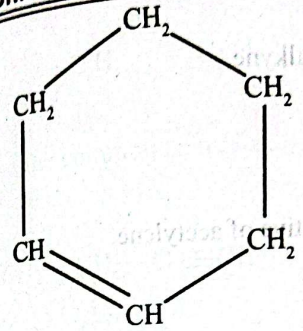
or 

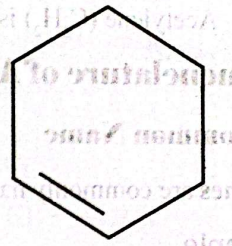
or C_4H_6 cyclobutene



or 

or C_5H_8 cyclopentene



or 

or C_6H_{10} cyclohexene

Illustrative Examples

63. Which of the following formula represents cyclic alkenes?

A) C_4H_8 B) C_4H_{10} C) C_4H_6 D) C_4H_9

Solution

In cycloalkenes, the number of hydrogen atoms is 2 times that of carbon minus two.

Answer: C

64. Which of the following is the molecular formula of cycloheptene?

A) C_7H_{12} B) $C_{17}H_{14}$ C) C_7H_{16} D) C_7H_{15}

Answer: A

65. Cylo alkenes are isomers with _____

Answer: Open chain alkynes

2. Alkynes

- They are unsaturated hydrocarbons
- They contain carbon - carbon triple bond.
- They have a general formula of C_nH_{2n-2}, $n \geq 2$

Example

For n = 2, C_2H_2

For n = 3, C_3H_4

The alkynes homologous series is also called **acetylene series.**

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$C_2H_2, C_3H_4, C_4H_6, C_5H_8$ -----

Acetylene (C_2H_2) is the simplest alkyne

Nomenclature of Alkynes

A) Common Name

Alkynes are commonly named as derivative of acetylene.

Example

$H-C\equiv C-H$ **Acetylene**

$H-C\equiv C-CH_3$ **Methyl acetylene**

$CH_3-C\equiv C-CH_3$ **Dimethyl acetylene**

$CH_3-CH_2-C\equiv C-CH_3$ **Ethyl methyl acetylene**

B) IUPAC Name of Alkynes

- The suffix added to the base name is " - yne"
- Start from the end nearer to the triple bond. The rest of the rules are similar with that of alkene.

Example

$HC\equiv CH$ **Ethyne**

$CH_3-C\equiv CH$ **Propyne**

$CH_3-C\equiv C-CH_3$ **-1-butyne or but -1- yne**

$^{4}CH_3-\overset{CH_3}{\overset{|}{^{3}CH}}-{}^{2}C\equiv {}^{1}CH$ **3 - methyl -1- butyne or 3 - methyl but -1- yne**

Illustrative Examples

66. The molecular formula of alkynes is

A) C_nH_{2n} B) C_nH_{2n-2} C) C_nH_{2n+2} D) C_nH_{2n+1}

Answer: B

67. The molecular formula of an alkyne with 15 carbon atoms is

A) $C_{15}H_{28}$ B) $C_{15}H_{30}$ C) $C_{15}H_{32}$ D) $C_{15}H_{31}$

Solution

$C_nH_{2n-2} = C_{15}H_{2\times15-2} = C_{15}H_{28}$

Answer: A

68. Give the common name for:

a. $CH_3-\overset{CH_3}{\overset{|}{CH}}-C\equiv CH$

b. $CH_3-\overset{CH_3}{\overset{|}{CH}}-C\equiv C-C_2H_5$

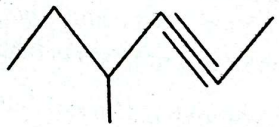
Solution

a) Isopropyl acetylene

b) Ethyl isopropyl acetylene

69. Give the IUPAC name of

a. $H_3C-C\equiv C-\underset{CH_3}{\underset{|}{CH}}-\underset{CH_3}{\underset{|}{CH}}-\underset{Br}{\underset{|}{C}}(-CH_2-CH_3)-CH_2-CH_3$

b. 

Solution

a) 6 - Bromo - 6 - ethyl - 4.5 - dimethyl - 2 - octyne

(6 - Bromo - 6 - ethyl - 4.5 - dimethyl oct - 2 - yne)

b) 4 - methyl - 2 - hexyne or 2 - methyl hex - 2 - yne

70. Give the IUPAC name for:

$CH_3-CH=CH-C\equiv CH$

Solution

If the double bond and triple bond appear in the same structure start counting from the end nearer to the triple bond.

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$CH_3-CH=CH-C\equiv CH$ (5 4 3 2 1) 3 - penten - 1 - yne

71. Draw the structure of:

a) 1 - chloro - 2 - butyne

b) 4 - methyl - 2 - pentyne

Solution

a. $ClCH_2-C\equiv C-CH_3$

b. $CH_3-C\equiv C-CH(CH_3)-CH_3$

Isomerism in Alkynes

- **Alkynes** exhibit structural and position isomerism
- Ethyne and propyne do not show any isomerism.

Example

Write the position isomers of butyne

Solution

$HC\equiv C-CH_2-CH_3$ **1- butyne**

$H_3C-C\equiv C-CH_3$ **2- butyne**

Physical Properties of Alkynes

- C_2 - C_4 are gases
- C_5 - C_{17} are liquids
- Above C_{17} are solids
- Alkynes are insoluble in water because they are non - polar.
- As carbon number increases boiling point, melting point and density increase.

Illustrative Examples

72. Which of the following alkynes does not have isomers?

A) Hexyne B) Octyne C) Propyne D) Pentyne

Answer: C

73. Which of the following has the highest boiling point?

A) $C_{12}H_{22}$ B) C_6H_{10} C) C_7H_{12} D) C_8H_{16}

Answer: A

74. Which one of the following has least boiling point?

A) 1 - Hexyne

B) 2 - Hexyne

C) 3 - methyl - 1 - pentyne

D) 3,3 -dimethyl - 1 - butyne

Solution

As branching increases, boiling point decreases.

Answer: D

75. Draw the position isomers of heptyne.

Solution

$HC\equiv C-CH_2-CH_2-CH_2-CH_2-CH_3$ **1- Heptyne**

$H_3C-C\equiv C-CH_2-CH_2-CH_2-CH_3$ **2 - Heptyne**

$H_3C-CH_2-C\equiv C-CH_2-CH_2-CH_3$ **3 - Heptyne**

Chemical Properties of Alkynes

Alkynes are the most reactive aliphatic hydrocarbons due to the presence of carbon - carbon triple bond.

1. **Addition reaction** is the common reaction of alkynes.

a) Addition of Hydrogen

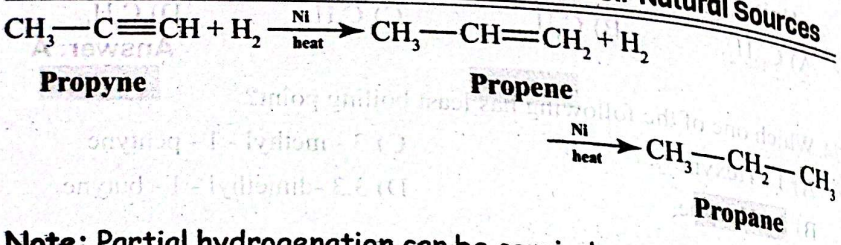
Example

$R-C\equiv CH + H_2 \xrightarrow[heat]{Ni} R-CH=CH_2 + H_2$

Alkyne **Alkene**

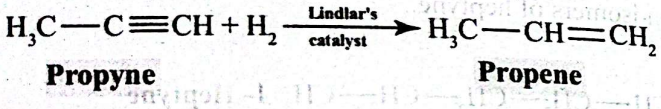
$\xrightarrow[heat]{Ni} R-CH_2-CH_3$

Alkane

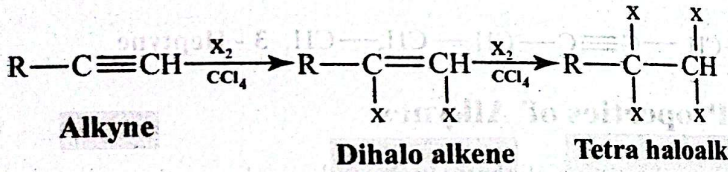


Note: Partial hydrogenation can be corried out by using a special called **Lindlar catalyst.**

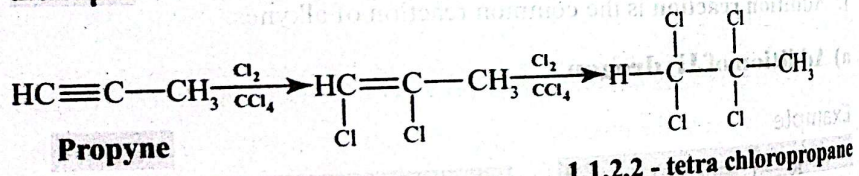
Example



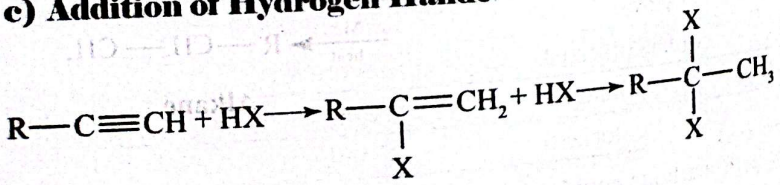
b) Addition of Halogen



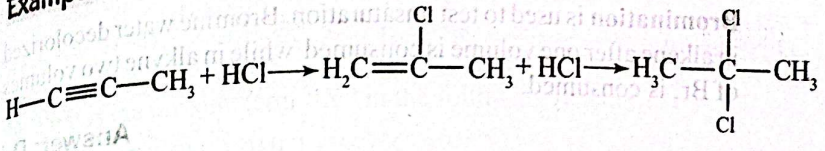
Example



c) Addition of Hydrogen Halide

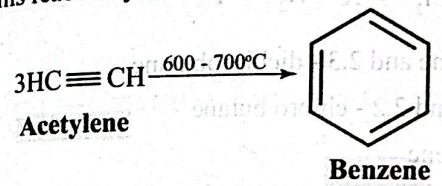


Example



d) Trimerization of Acetylene

☑ This reaction yields benzene



2. Combustion of Alkynes

Alkynes burn with smoky luminous flame and release greater quantity of heat than alkenes of similar carbon number due to carbon - carbon triple bond.

Alkynes + oxygen (excess) → carbondioxide + water

Alkynes + oxygen (limited) → carbonmonoxide + water

Example

$2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$

Illustrative Examples

76. Identification of the unknown hydrocarbons of equal number of moles has shown the following experimental results in the laboratory.

Hydrocarbon	Addition of Br_2 water	Amount of Br_2 consumed
A	Decolorized	1 Volume
B	No change	None
C	Decolorized	2 Volumes

The three hydrocarbons could respectively be

A) Alkane, alkene, alkyne
B) Alkene, alkyne, alkane
C) Alkene, alkyne, alkane
D) Alkene, alkane, alkyne

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Solution

Bromination is used to test unsaturation. Bromine water decolorized in alkene after one volume is consumed while in alkyne two volumes of Br_2 is consumed.

Answer: D

77. What are the products 'X' and 'y' in the reaction below respectively?

$CH_3—C\equiv C—CH_3 + Cl_2 \longrightarrow X + H_2 \longrightarrow Y$

A) 2,3 - dichloro 2 - butene and 2,3 - dichlorobutane

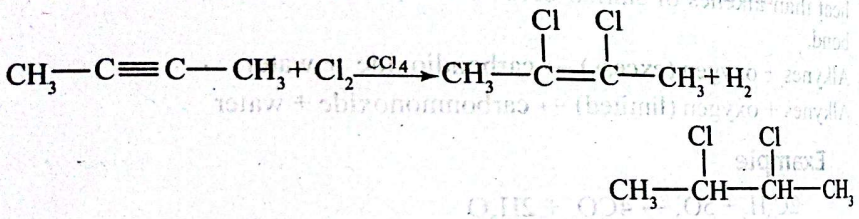
B) 2,2 - dichlorobutene and 2,2 - chloro butane

C) 2 - butyne and 2 - butene

D) 1,2 - dichloro butene and 1,2 - dichlorobutane.

Solution

2,3 - dichloro - 2 - butene



2, 3 - dichloro butane

Answer: A

78. When one mole of acetylene is completely reacted in excess amount of bromine. Which product is likely to occur?

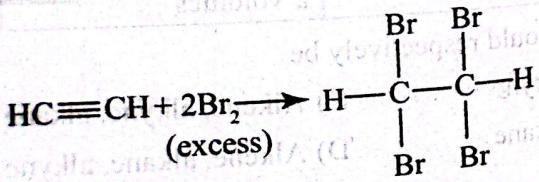
A. $CH_3—CHBr_2$

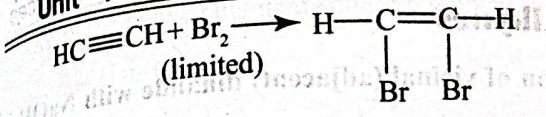
B. $CHBr_2—CHBr_2$

C. $CHBr=CHBr$

D. $CH_3—CH_2Br$

Solution



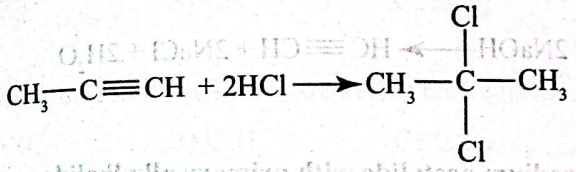


Answer: B

79. What is the missing term "X" in the following reaction?

$CH_3—C\equiv CH + 2HCl \longrightarrow X$

Solution



80. Complete and balance the following

a) $C_2H_2 + 2H_2 \longrightarrow$

b) $C_2H_2 + 2Cl_2 \longrightarrow$

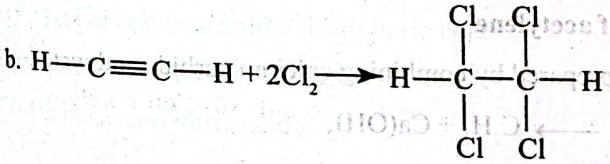
c) $C_2H_2 + 2HCl \longrightarrow$

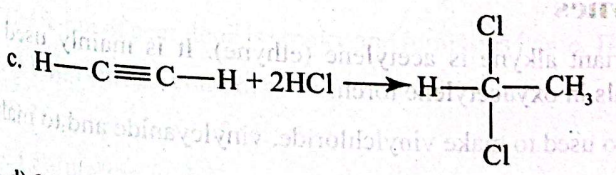
d) $C_2H_2 + O_2 \longrightarrow$

Solution

a) $C_2H_2 + 2H_2 \longrightarrow C_2H_6$

Ethyne Ethane





d) $2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O$

Preparation of Alkynes

1. **Dehydrohalogenation of vicinal (adjacent) dihalide with NaOH or KOH or $NaNH_2$.**

$$R-\underset{X}{CH}-\underset{X}{CH_2} + 2NaOH \longrightarrow \underset{\text{Alkyne}}{R-C\equiv CH} + 2NaX + 2H_2O$$

Example

$$\underset{Cl}{CH_2}-\underset{Cl}{CH_2} + 2NaOH \longrightarrow HC\equiv CH + 2NaCl + 2H_2O$$

2. **Alkylation of sodium acetylide with primary alkylhalide**

$$HC\equiv C^{-}Na^{+} + R-X \longrightarrow \underset{\text{Alkyne}}{HC\equiv C-R} + NaX$$

Sodium carbide **1° alkyl halide**

Example

$$HC\equiv C^{-}Na^{+} + CH_3Cl \longrightarrow HC\equiv C-CH_3 + NaCl$$

3. **Production of acetylene**

Acetylene is prepared by combining calcium carbide and water.

$$CaC_2 + 2H_2O \longrightarrow C_2H_2 + Ca(OH)_2$$

Uses of Alkynes

The most important alkyne is acetylene (ethyne). It is mainly used for cutting and metals in oxyacetylene torch.

Acetylene is also used to make vinylchloride, vinylcyanide and to make a polymer acrylon.

Illustrative Examples

81. A reaction for the general preparation for a member of homologous series of certain hydrocarbon is given by:

$$R-CHCl-CH_2Cl + 2NaOH \longrightarrow X + 2NaOH + 2H_2O$$

"X" in the above chemical equation represents

A. $R-C\equiv CH$

B. $R-\underset{CH_3}{CH}-CH_3$

C. $R-CH=CH_2$

D. $R-CH_2-CH_3$

Answer: A

82. Given the reaction: $CaC_2 + 2H_2O \longrightarrow X + Ca(OH)_2$, X is

A) C_2H_2 B) CH_4 C) C_2H_4 D) C_2H_6

Answer: A

83. Which of the following hydrocarbons has the highest heat of combustion?

A) CH_4 B) C_2H_6 C) C_2H_4 D) C_2H_2

Solution

Acetylene (C_2H_2) has the highest heat of combustion due to the carbon - carbon triple bond.

Answer: D

84. Which of the following reagents are used to prepare acetylene in the laboratory?

A) Sodium acetate and soda lime

B) Alcohol and concentrated sulphuric acid

C) Calcium carbide and water

D) Benzene and nitric acid

Answer: C

85. The flame of acetylene is smoky and luminous flame. This is because of:

A) High carbon content of the compound

B) High reactivity of acetylene

C) Saturation of the compound

D) Low molecular mass

Answer: A

86. Complete and balance the following reaction:

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a) $CH_3—CHBr—CHBr—CH_3 + 2KOH \longrightarrow$

b) $H—C\equiv C^-Na^+ + CH_3—\underset{\displaystyle CH_3}{\underset{|}{CH}}—CH_2Cl$

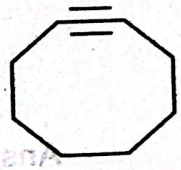
Solution

a) $CH_3—CHBr—CHBr—CH_3 + 2KOH \longrightarrow CH_3—C\equiv C—CH_3 + 2KBr + 2H_2O$

b) $H—C\equiv C^-Na^+ + CH_3—\underset{\displaystyle CH_3}{\underset{|}{CH}}—CH_2Cl \longrightarrow H—C\equiv C—CH_2—\underset{\displaystyle CH_3}{\underset{|}{CH}}—CH_3$

Cycloalkynes

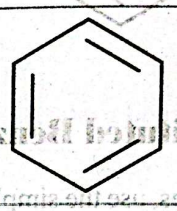
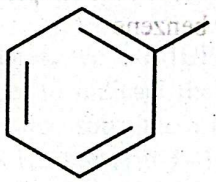
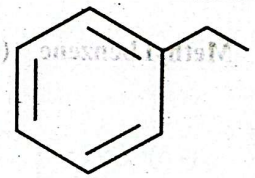
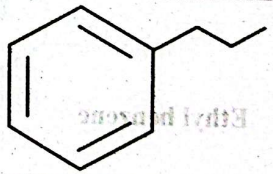
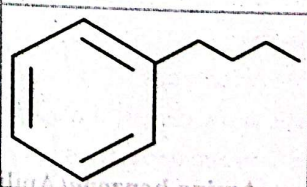
- ☑ They are cyclic alkynes
- ☑ Their general formula is C_nH_{2n-4}, $n \geq 8$, because of the linear nature of the $-C\equiv C-$ alkyne of unit, cycloalkynes are usually highly strained and can only exist when the number of carbon atoms in the ring is greate enough to provide the flexibility necessary to accommodates this geometry.
- ☑ **Cyclo octyne** is the simplest cycloalkyne that is stable compound.

C_8H_{12}  Cyclooctyne

4. Aromatic Hydrocarbons

- ➥ They are unsaturated hydrocarbons
- ➥ They are stable towards many chemical reagents.
- ➥ They are benzene and benzene derivatives.
- ➥ They have a general formula of C_nH_{2n-6}, $n \geq 6$

Table 11.6 Some Aromatic Hydrocarbons, their Structural and Molecular Formula.

Aromatic Hydrocarbons	Structural Formula	Molecular Formula
Benzene		C_6H_6
Toluene (methyl benzene)		C_7H_8
Ethyl benzene		C_8H_{10}
Propyl benzene		C_9H_{12}
Butyl benzene		$C_{10}H_{14}$

- ☑ **Benzene** is the simplest aromatic hydrocarbon.
- ☑ Its structure was determined by **F. Kekule.**

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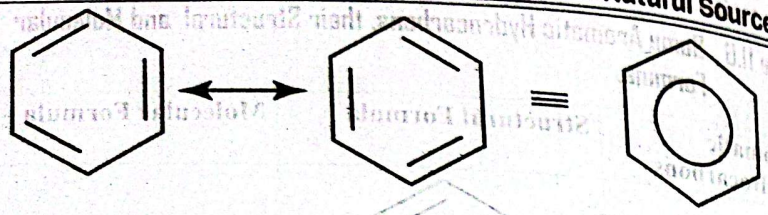
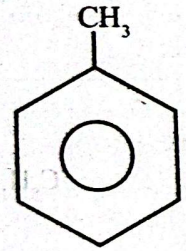


Resonance hybrid

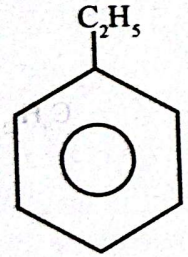
Nomeclature of Substituted Benzenes

To name monosubstited benzenes, use the simple method below. The name of the side chain following by benzene

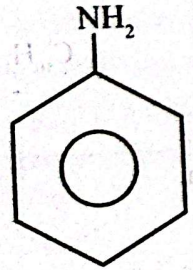
Example

a) 

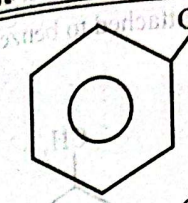
Methyl benzene (toluene)common name

b) 

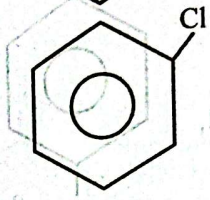
Ethyl benzene

c) 

Amino benzene(Aniline) common name

d) 

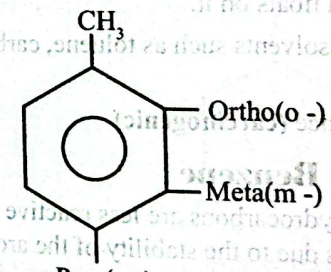
Hydroxy benzene (Phenol) common name

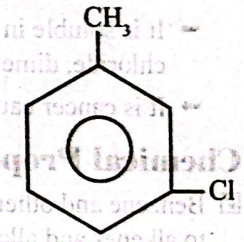
e) 

Chloro benzene

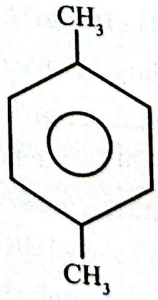
For **disubstituted benzenes,** we use IUPAC and common names. For IUPAC name, use number to indicate the position of side chain in the nearest direction to the mono substituted benzenes. For common name, used prefixes ortho (o -) - 1,2, meta (m-) - 1,3 and para (p-) (1 - 4)

Example

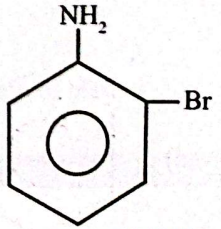




3 - chloro toluene
(m- chloro toluene)



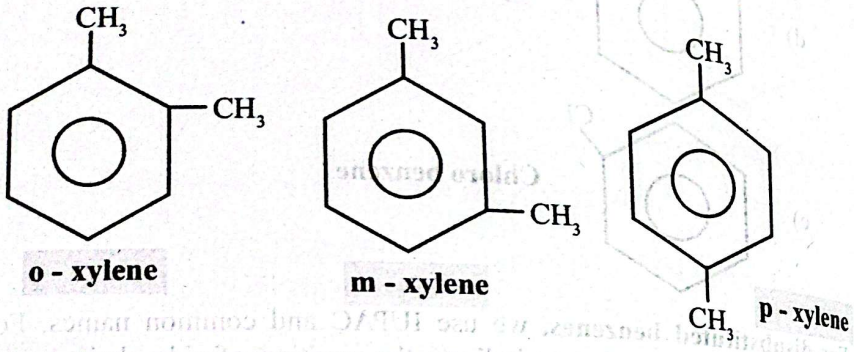
4 - methyl toluene (p - xylene)



2 - Bromo aniline
(o - Bromo aniline)

Xylene is a compound with two methyl groups attached to benzene ring.

Example



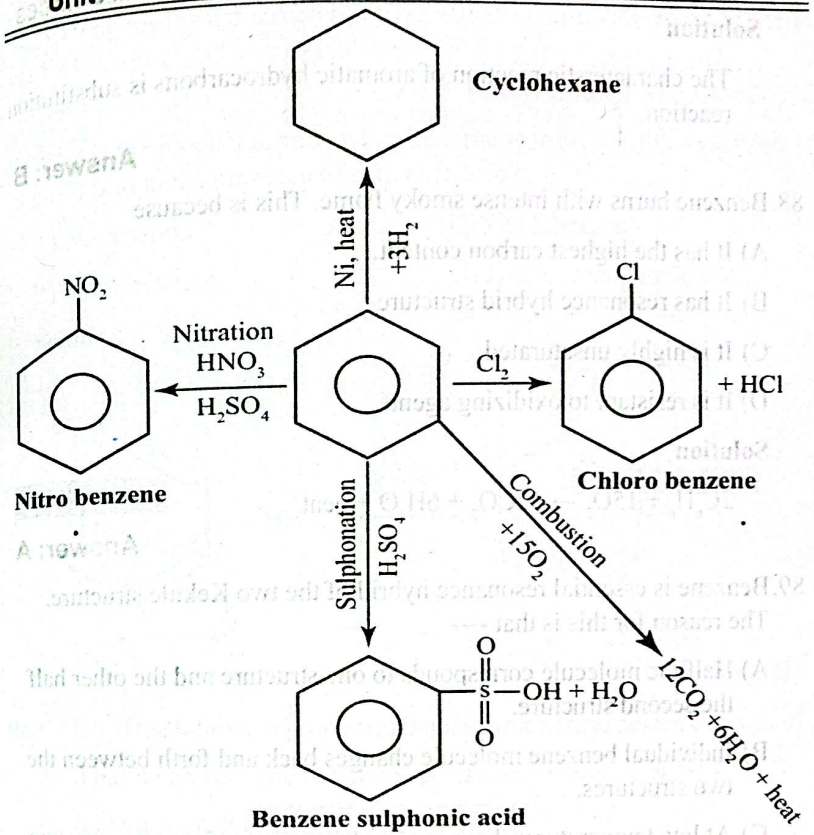
Physical Properties of Benzene

- It is colorless, flammable and volatile liquid
- It has a characterstic smell.
- It is insoluble in water and floats on it.
- It is soluble in non - polar solvents such as toluene, carbontatra chloride, dimethylether
- It is cancer causing substance **(carcinogenic)**

Chemical Properties of Benzene

☑ Benzene and other aromatic hydrocarbons are less reactive compared to alkenes and alkynes. This is due to the stability of the aromatic ring.

☑ The characterstic reaction of benzene is **substitution reaction.**



Note: Nitration, sulphonation and halogenation are substitution reactions.

Uses of Aromatic Hydrocarbons

They are used as organic solvents

Benzene is used as fuel, raw materials for production of plastics, materials and other benzene derivatives.

Illustrative Examples

87. Aromatic hydrocarbons mainly react by

A) Addition C) Elimination

B) Substitution D) Hydrogenation

Solution

The characterstic reaction of aromatic hydrocarbons is substitution reaction.

Answer: B

88. Benzene burns with intense smoky flame. This is because

A) It has the highest carbon content.

B) It has resonance hybrid structure

C) It is highly unsaturated

D) It is resistant to oxidizing agents.

Solution

$$2C_6H_6 + 15O_2 \rightarrow 12CO_2 + 6H_2O + \text{heat}$$

Answer: A

89. Benzene is essential resonance hybrid of the two Kekule structure. The reason for this is that ----

A) Half the molecule corresponds to one structure and the other half the second structure.

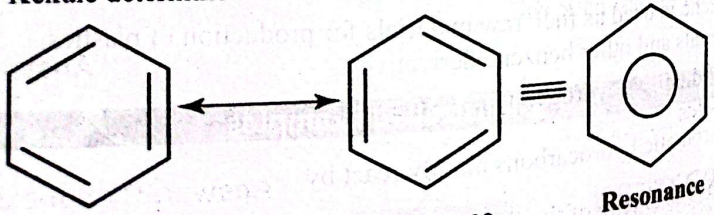
B) Individual benzene molecule changes back and forth between the two structures.

C) At low temperatures, benzene can be separated into two structure.

D) The two structures make equal contributions to the resonance hybrid.

Solution

Kekule determined the structure of benzene.



Resonance Structure I **Resonance Structure II** **Resonance Hybrid**

Both resonance structures have equal contribution to the hybrid structures.

Answer: D

90. What is the product formed when benzene is treated with concentrated nitric acid in the presence of sulphuric acid?

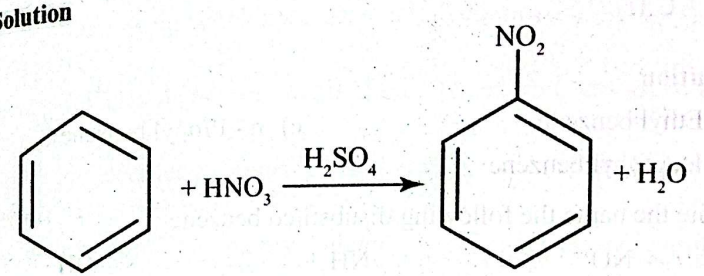
A) No reaction

B) Nitro benzene

C) Cyclohexane

D) Benzene sulphonic acid

Solution



Nitro benzene

Answer: B

91. Which of the following properties is not appropriate to describe benzene?

A) It has a characterstic reactivity of alkenes

B) It has a resonance hybrid structure

C) It is an aromatic hydrocarbon

D) It is volatile liquid.

Solution

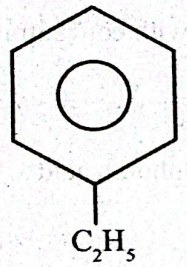
Benzene is less reactive than alkenes and alkynes due to the stability of the aromatic ring

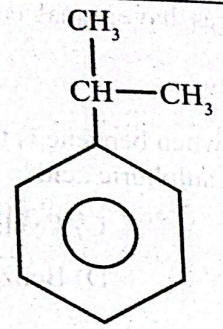
Answer: A

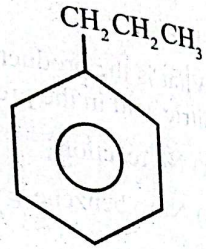
92. Addition of hydrogen to benzene in the presence of catalyst and heat yields _____

Answer: Cyclohexane.

93. Give the name of the following mono substituted benzene

a) 

b) 

c) 

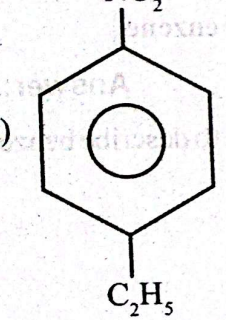
Solution

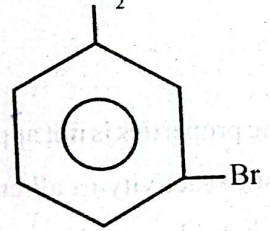
a) Ethyl benzene

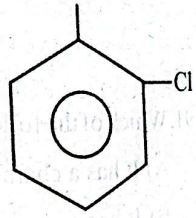
b) Isopropyl benzene

c) n - Propyl benzene

94. Write the name the following disubstited benzenes.

a) 

b) 

c) 

Solution

a) 4 - Ethyl nitro benzene (p - ethyl nitro benzene)

b) 3 - Bromo amino benzene (m - Bromoamino benzene)

c) 2 - Chloro phenol (o - chloro phenol)

11.4 Natural Sources of Hydrocarbons

The natural sources of hydrocarbons are natural gas, petroleum and coal.

1. **Natural gas** - mainly contains the first four alkanes methane, ethane, propane and butane.

Natural gas is a fossil fuel formed from ancient plants and animals buried under layer of rock and sand

Uses of Natural Gas

- Heating, cooling and cooking in residential, commercial and industrial settings.
- Fueling vehicles that run on natural gas which emit less green house gases than gasoline.
- Generating electricity in power plants that use natural gas as a source energy.

2. **Petroleum** contains mainly alkanes, cycloalkanes and aromatic compounds.
- It cannot be used for various purposes directly.
- **Fractional distillation** changes petroleum to useful products. This physical technique separates mixtures that contain different components with different **boiling point**.
- **Cracking** is the most important process that breaks long chain hydrocarbons into smaller units.

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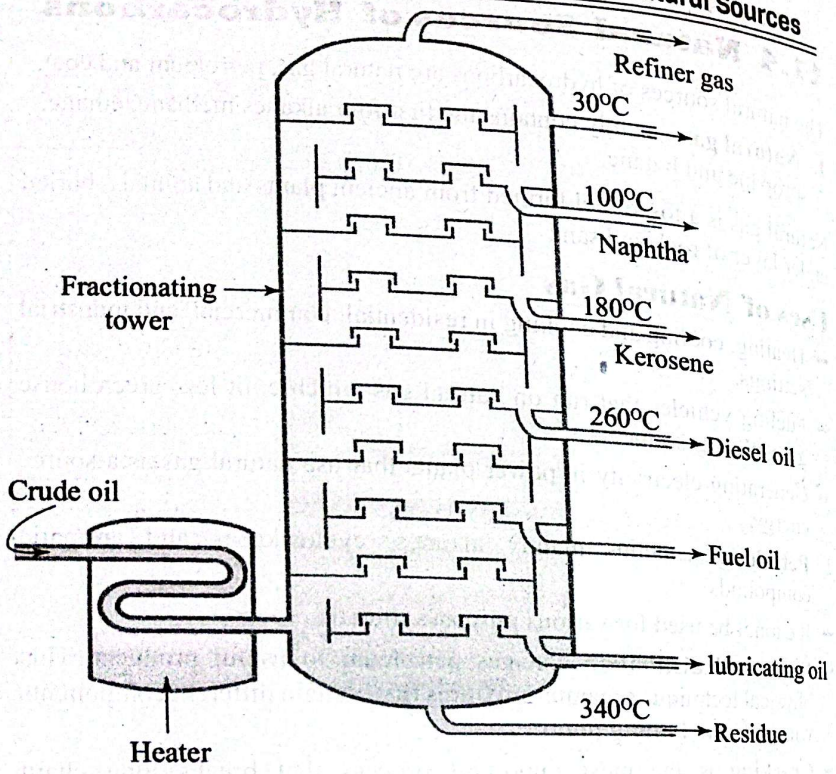

Fig 11. 2 The different Fractions Obtained from Crude Oil

Table 11.7 Common Petroleum Products

Fraction	Approximate carbon number	Boiling Point (°C)	Uses
Gases	C_1 - C_4	Below 20	➥ As fuel in the form of bottled gas
Petroleum ether	C_5 - C_7	20 - 60	➥ As solvent, in dry cleaning
Gasoline (petrol)	C_5 - C_{10}	40 - 200	➥ Motor fuel for internal combustion engines
Kerosene	C_{10} - C_{16}	175 - 275	➥ Jet engine fuel ➥ House hold fuel
Fuel oil (Diesel oil)	C_{15} - C_{18}	250 - 350	➥ Furnace fuel, diesel engine fuel
Lubricating greases petroleum jelly	> C_{19}	Non - volatile	➥ Lubrication
Paraffin Wax Asphalt (Bitumen) coke	> C_{20}	Non - volatile	➥ Candles, polishes. ➥ Roofing and road construction ➥ Electrodes, fuel for power stations.

3. **Coal** is a compbustible black or dark brown rock consisting chiefly of carbonized plant matter, found mainly in under ground seas and used as fuel.

➥ It is an important sources of aromatic hydrocarbons.

➥ **Destructive distillation of coal** is heating coal in the absence of oxygen. This results volatile products and coal.

Volatile products → Coal gas + Coal tar (liquid)

➥ **Coke** is pure and solid form of coal which is used as fuel to produce gaseous fuels.

Water gas = $CO + H_2$

Producer gas = $CO + N_2$

Table 11.8 The Different Coal Types

Type of coal	Carbon content (%)
Anthracite	90
Bituminous coal	60
Lignite	20
Peat	

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Illustrative Examples

95. Which physical property makes it possible to separate the components of crude oil by means of fractional distillation?

A) Boiling point
B) Melting point
C) Conductivity
D) Density

Answer: A

96. Water gas a mixture of

A) CO and H_2O
B) CO and H_2
C) H_2 and CH_4
D) CO and N_2

Solution

Water gas = CO + H_2 in one to one mole ratio.

Answer: B

97. Which of the following products of petroleum are arranged in correct order of increasing boiling point?

A) Petrol, asphalt, diesel oil, kerosene
B) Asphalt, diesel oil, kerosene, petrol
C) Diesel oil, petrol, asphalt, kerosene
D) Petrol, kerosene, diesel oil, asphalt.

Solution

- Gases C_1- C_4
- Petroleum ether C_5 - C_{17}
- Gasoline C_5 - C_{10}
- Kerosene C_{10} - C_{16}
- Fuel oil (desel oil) C_{10} - C_{18}
- Ludricating / Grease } > C_{19}
- Paraffin wax / Asphal coke } > C_{20}

Answer: D

98. The main constituent of petroleum are

A) Alkenes and natural gas
B) Alkenes and cycloalkanes
C) Alkanes and aromatics
D) Benzene and alkenes

Answer: B

99. A petroleum product commonly used for jet engines and tractors, heating and lighting purpose is

A) Petrol
B) Petrolum
C) Kerosene
D) Diesel oil

Answer: C

100. Which of the following substances formed from the remains of plants and animals?

A) Natural gas
B) Petroleum
C) Coal
D) Water gas

Solution

Natural gas originated from the decay and decomposition of plants and animals?

Answer: A

101. Why is crude oil unstable when it is extracted?

Solution

Because it is a mixture of compounds that must first be refined.

102. The general process of converting the mixture that is collected as crude oil into separate fractions is _____.

Answer: Oil refining or petroleum refining

103. Crude oil is a mixture of hydrocarbons. The refining of crude oil produces fractions which are more usefull to us than crude oil itself. The separation is carried out in fracting column. As shown below.

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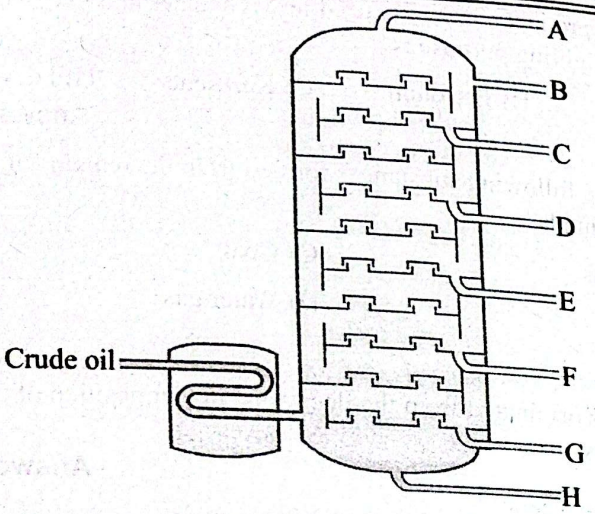


a) Which separation technique is used to separate the fractions?

b) Name each of the fractions A to H

Solution

a) Fractional distillation

b) Refinery gas - used as fuel

c) Gasoline - used as fuel in cars (petrol)

d) Naphtha - used to make chemicals

e) kerosene /paraffin - used as fuel in jet engines

f) Diesel oil or gas oil - used as fuel in diesel engines

g) Fuel oil - used as a fuel for ships and for home heating systems

h) Lubricating oil - Lubrication

i) Residue - used to make bitumen for surfacing roads.

Solved problems

104. For hexane, write its

a) Molecular formula

b) Condensed formula

c) Detail structural formula

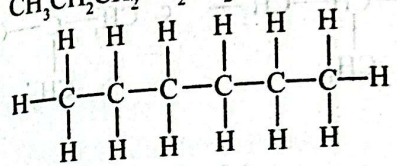
d) Line formula

Solution

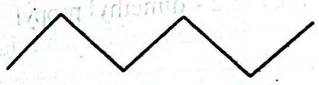
a) C_6H_{14}

b) $CH_3CH_2CH_2CH_2CH_2CH_3$ or $CH_3(CH_2)_4CH_3$

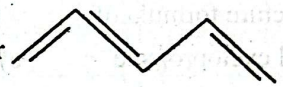
c)



d)



105. Give the IUPAC name for



Solution

1,2,4 - pentatriene

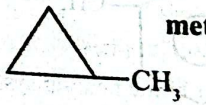
106. Write structural formula for all structural isomers of cyclobutane.

Solution

$CH_2{=}CH{-}CH_2{-}CH_3$ **1 - butene**

$CH_3{-}CH{=}CH{-}CH_3$ **2 - butene**

$CH_3{-}C(CH_3){=}CH_2$ **2 - methyl propene**



methyl cyclopropane

107. Write the structural formula for

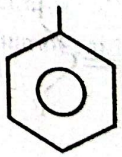
a) Vinyl b) Allyl c) Phenyl d) Benzyl

Solution

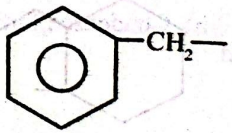
a. $CH_2{=}CH{-}$

b. $CH_2{-}CH{=}CH{-}$

c. $C_6H_5{-}$ or



d. $C_7H_7{-}$ or



108. Give IUPAC name of the following alkyl groups.

a) $CH_3-CH(CH_3)-$ b. $CH_3-CH(CH_3)-CH_2-$ c. $CH_3-C(CH_3)(CH_3)-CH_2-$

Solution

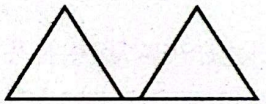
a) 1 - Methylethyl
b) 2 - Methyl propyl
c) 2,2 - dimethyl propyl

109. Draw the structure formula of:

a) Cyclopropyl cyclopropane
b) 1,1 - dichloro - 1- butene
c) o - Bromo aniline
d) o - ethyl - m - mehyl phenol
e) 3,3 - dimethyl hex - 1 - yne
f) m - xylene
g) Cyclopentene
h) Tetrafluoro ethylene

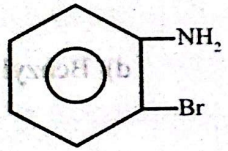
Solution

a.

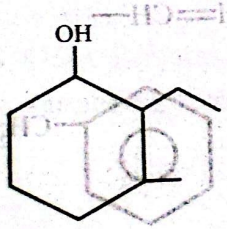


b. $Cl_2C=CH-CH_2-CH_3$

c.



d.



e. $HC\equiv C-C(CH_3)(CH_3)-CH_2-CH_2-CH_3$

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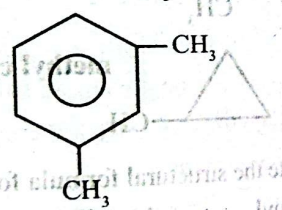

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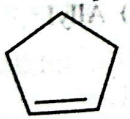

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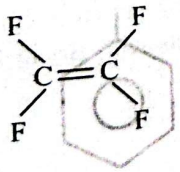
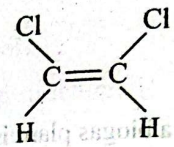
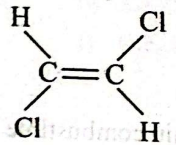


110. Draw the cis - trans isomers of 1,2 - dichloroethene

Solution



cis - 1,2 - dichloro ethene



Trans - 1,2 - dichloro ethene

111. Give the following table

Component of natural gas	Boiling point in °C
Methane	-161. 6
Ethane	- 88.62
Propane	- 41 . 1
Butane	- 0.5

a) What physical technique is used to separate the component of natural gas?
b) Arrange the components of natural gas increasing order of strength of London force.

Solution

a) Fractional distillation
b) Methane < ethane < propane < butane

112. Propyne can be prepared using the following equation: $X + 2NaOH \longrightarrow CH_3-C\equiv CH + 2NaBr + 2H_2O$. What is the missing reactant X ?

Solution

$CH_3-CH(Br)-CH_2(Br)$ **1,2 - dibromopropane**

113. Benzene reacts with Br_2 in the presence of $FeCl_3$, Write the major product formed.

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Solution

$$C_6H_6 + Br_2 \xrightarrow{FeCl_3} C_6H_5Br + HBr$$

Bromo benzene

Major product

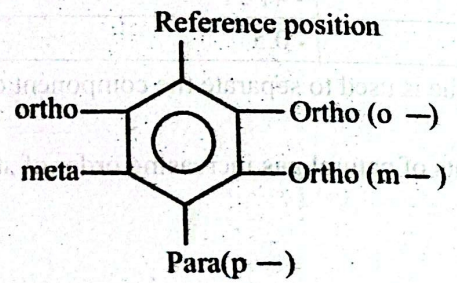
114. The main combustible gas obtained from a biogas plant is ____.

Solution

The biogas plant operates on the simple principle that when dung or any other organic matter is fermented in the absence of air. The combustible gas produced is methane (CH_4).

115. Represent the ortho, meta and para substituted benzene.

Solution



116. Classify the following organic compounds according to their functional groups.

a. $H_2C{=}CH{-}CH_3$

b. $H_3C{-}O{-}CH_3$

c. $CH_3{=}CH_2{-}CH_2{-}$

d. $H_3C{-}\overset{\overset{O}{\|}}{C}{-}O{-}CH_2{-}CH_3$

e. $CH_3{-}CH_2{-}CH_2{-}\overset{\overset{O}{\|}}{C}{-}OH$

f. $CH_3{-}\overset{\overset{O}{\|}}{C}{-}CH_2{-}CH_3$

g. $CH_3{-}CH_2{-}\overset{\overset{O}{\|}}{C}{-}H$

h. $CH_3{-}C{\equiv}C{-}CH_3$

i. $CH_3{-}CH_2{-}CH_3$

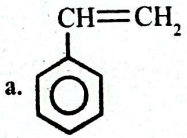
Solution

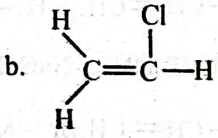
a) Alkene
b) Ether
c) Alcohol
d) Ester
e) Carboxylic acid
f) Ketone
g) Aldehyde
h) Alkyne
i) Alkane

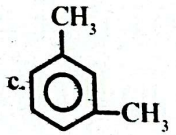
117. Draw the structure of

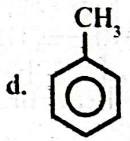
a) Styrene
b) Vinyl chloride
c) m - xylene
d) Toluene
e) Phenol
f) Nitro benzene
g) o - dibromobenzene
h) 1,3,5 - trichlorobenzene
i) 1,2,4 - triethyl benzene

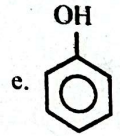
Solution

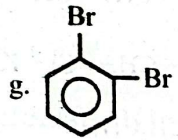
a. 

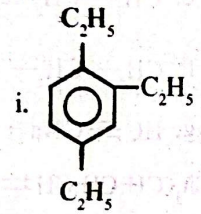
b. 

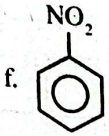
c. 

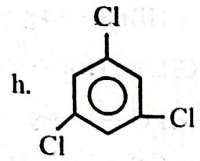
d. 

e. 

g. 

i. 

f. 

h. 

118. Starting with the appropriate alkene, tell how you could prepare each of the following by a single reaction

a) 1,2 - Dibromobutane
b) 2 - Bromo butane
c) 2 - Butanol

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Solution

a. $CH_3-CH_2-CH=CH_2 + Br_2 \xrightarrow{CCl_4} CH_3-CH_2-\underset{Br}{CH}-\underset{Br}{CH}$

b. $CH_3-CH_2-CH=CH_2 + HBr \xrightarrow{CCl_4} CH_3-CH_2-\underset{Br}{CH}-CH_3$

c. $CH_3-CH_2-CH=CH_2 + H_2O \xrightarrow{H^+} CH_3-CH_2-\underset{OH}{CH}-CH_3$

119. Complete and balance the following reactions

a. $CH_3-CH=CH_2 + H_2 \xrightarrow{Ni}$

b. $CH_3CH_2COONa + NaOH \xrightarrow[heat]{CaO}$

c. $CH_3-CH_2=CH_2Br + Na \longrightarrow$

d. $CH_3-CH_2-\underset{Br}{CH}-\underset{Br}{CH_2} + NaOH \longrightarrow$

e. $CH_3-CH_2-CH_2-CH_2-OH \xrightarrow[heat]{H_2SO_4}$

f. $CH_3-CH_2-CH_2-CH_2-Cl + KOH \xrightarrow[heat]{H_2SO_4}$

g. $HC\equiv CNa + CH_3CH_2CH_2Br \longrightarrow$

h. $CH_3CH_2CH=CH_2 + HBr \longrightarrow$

g. $HC\equiv C-CH_2-CH_3 + 2HBr \longrightarrow$

j. $CH_3-CH=CH_2 + H_2O \xrightarrow{H^+}$

j. $CH_3-CH_2-CH_2-CH=CH_2 + H_2O \xrightarrow[KMnO_4]{Alkaline}$

Solution

a. $CH_3-CH=CH_2 + H_2 \xrightarrow[heat]{Ni} CH_3-CH_2-CH_3$

b. $CH_3CH_2COONa + NaOH \xrightarrow[heat]{CaO} CH_3-CH_3 + Na_2CO_3$

c. $CH_3-CH_2-CH_2Br + 2Na \longrightarrow CH_3-CH_2-CH_2-CH_2-CH_2-CH_3 + 2NaBr$

d. $CH_3-CH_2-\underset{Br}{CH}-\underset{Br}{CH_2} + NaOH \longrightarrow CH_3-CH_2-C\equiv CH + 2NaBr + 2H_2O$

e. $CH_3-CH_2-CH_2-CH_2-OH \xrightarrow[heat]{H_2SO_4} CH_3-CH_2-CH=CH_2 + H_2O$

f. $CH_3-CH_2-CH_2-CH_2-Cl + KOH \xrightarrow[heat]{H_2SO_4} CH_3-CH_2-CH=CH_2 + KCl + H_2O$

g. $HC\equiv CNa + CH_3CH_2CH_2Br \longrightarrow HC\equiv CCH_2CH_2CH_3 + NaBr$

h. $CH_3CH_2CH=CH_2 + HBr \longrightarrow CH_3-CH_2-\underset{Br}{CH}-CH_3$

i. $HC\equiv C-CH_2-CH_3 + 2HBr \longrightarrow H_3C-\underset{Br}{\overset{Br}{C}}-CH_2-CH_3$

j. $CH_3-CH=CH_2 + H_2O \xrightarrow{H^+} CH_3-\underset{OH}{CH}-CH_3$

k. $CH_3-CH_2-CH_2-CH=CH_2 + H_2O \xrightarrow[KMnO_4]{Alkaline} CH_3-CH_2-CH_2-\underset{OH}{\overset{OH}{CH}}-CH_2$

120. Write the two reagents used to test unsaturation

Solution

Bromine water (Br_2 in CCl_4) and alkaline $KMnO_4$.

121. Write the most commonly used chemicals in dry cleaning.

Solution

- Tetra chloromethane, CCl_4
- Tetrachloroethylene, $Cl_2C = CCl_2$
- Benzene and gasoline

122. Write the agricultural applications of organic compounds.

Solution

To manufacture pesticides, herbicides and fertilizers.

123. Write the industrial applications of organic compounds.

Solution

Organic compounds are used to make alcoholic beverages, pharmaceticals, soaps and detergents, dry cleaning and fuels.

124. What are the natural sources of hydrocarbons?

Solution

Natural gas, petroleum and coal

125. What is the purpose of cracking during fractional distillation of crude oil?

Solution

It maximizes the out put of gasoline or petrol.

Supplementary Exercise

I. True/ False items

1. The first organic compound was synthesized in the laboratory by Friedrich Wohler.
2. 2,2 - dimethyl - 5 - ethylhexane is correct IUPAC name.
3. Addition reaction is the common reaction of alkenes and alkynes.

4. Hexene and cyclohexane are isomers.
5. Alkenes are the most reactive hydrocarbons.
6. There is aromatic hydrocarbon without benzene ring.
7. Organic compounds are classified based on the functional group they contain in their molecules.
8. Benzene mainly undergoes substitution reactions
9. The general formula of alkane homologous series is C_nH_{2n+1}
10. Producer gas is a mixture of carbon monoxide and nitrogen in one to one ratio

II. Multiple choice items

11. The simplest and first alkene is

A) Ethane B) Methane C) Ethene D) Ethyne

12. The characterstic group of an organic ester is

A. —CO— B. —COOH— C. —COO— D. —CHO—

13. Compounds that have the same composition but differ in their structural formula are

A) Isomers C) Substitution products

B) Polymers D) Alkanes

14. Coke is produced from bitmuminous coal by

A) Cracking C) Substitution

B) Synthesis D) Destructive distillation

15. Acid catalyzed hydration of ethylene gives

A) Ethyne B) Ethene C) Ethanol D) Ethanal

16. All of the following exhibit cis - trans isomers. EXCEPT

A) 1,2 - dichloroethene C) 1,1 - dibromopropene

B) 2 - butene D) 2 - penten

17. Examine the structures below, then indicate which one has the highest boiling point?

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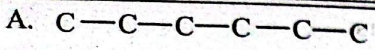
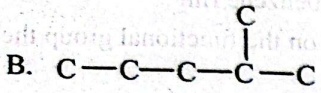
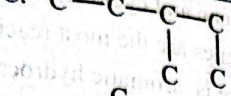
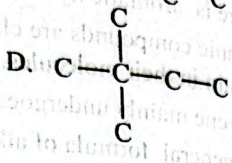
A. 

B. 

C. 

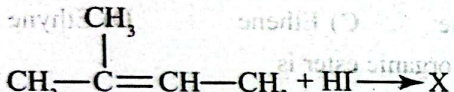
D. 

18. Which of the following react with hydrogen and nickel to form propane?

A) $CH_3CH_2CH_2OH$ C) $CH_3CHOHCH_3$

B) $CH_3CH = CH_2$ D) CH_3CH_2CHO

19. Consider the reaction below



: The IUPAC name of X is:

A) 2 - Iodo - 2 - methyl butane C) 2 - Iodo - 3 - methyl butane

B) 1 - Iodo - 2 - methyl butane D) 3 - Iodo - 2 - methyl butane

20. What do $CH_3 - CH_2 - CH_2 - CH_3$ and $(CH_3)_2 CH CH_3$ have exactly in common?

A) Melting point C) Chemical property

B) Boiling point D) Molecular mass

21. Which of the following elements is found in most organic compounds combined with carbon?

A) Nitrogen B) Hydrogen C) Oxygen D) Sulfur

22. Hydrocarbons with general formula C_nH_{2n-2} correponds to all of the following. **EXCEPT**

A) Alkynes C) Cycloalkenes

B) Dienes D) Cycloalkanes

23. Which of the following is liquid at root temperature?

A) C_4H_{10} B) C_2H_{40} C) C_2H_8 D) $C_{10}H_{20}$

24. All of the following is true about biogas. EXCEPT

A) It is cleaner and more readily available than other fuel resources..

B) Biogas usage would greatly reduce deforestation

C) Biogas technology has no negative effect on environment.

D) Biogas mainly contains ethane.

25. Which of the following mismatched with its definition?

A) Chain - propagating step: a reaction in which free radicals are consumed and generated

B) Photochemical reaction: reaction caused by the presence of light.

C) Free redical: atom or group of atoms possessing unpaired electrons

D) Combustion: endothermic reaction that forms carbon dioxide and water.

26. Which of the following is not true about benzene?

A) It is liquid at room temperature

B) It is colorless volatile liquid

C) It is polar compound

D) It is less dense than water

27. Compared to inorganic compounds, organic compounds:

A) have higher melting point

B) are more readily decomposed by heat

C) make very conductive water solution

D) are highly polar

28. Benzene belongs to a class of hydrocarbons called

A) Paraffins C) Aromatics

B) Olefins D) Acetylenes

29. Ethyne decolorizes Br_2 in CCl_4. This is due to the formation.

A) 1,1 - dibromethane C) 1,1,2,2 - tetra bromo ethane

B) 1,2 dibromoethane D) 1,1,2,2 - tetra bromo ethene

30. Which of the following hydrocarbon does not belong to the homologous series of alkynes?

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A) $C_{18}H_{34}$ B) $C_{15}H_{30}$ C) $C_{16}H_{30}$ D) $C_{20}H_{38}$

31. The vital force theory was proposed by

A) Wohler B) Berzelius C) Kolbe D) Dalton

32. Which of the following compounds contains maximum number of tertairy hydrogen atoms?

A) n - Hexane C) 2,2 - Dimethyl butane

B) 2,3 - Dimethyl butane D) 2 - Methyl pentane

33. Which of the following is the most stable?

A. $CH_3-\underset{\displaystyle CH_3}{\underset{|}{C}}=CH-CH_2-CH_3$ C. $CH_3-CH=CH-CH_3$

B. $CH_2=CH_2$ D. $CH_3-CH_2-CH=CH_3$

34. Which of the following represents the sec - butyl radical?

A. $CH_3-CH_2-CH_2-CH_2-$ C. $CH_3-\underset{\displaystyle CH_3}{\underset{|}{CH}}-CH_2-$

B. $CH_3-\underset{|}{CH}-CH_2-CH_3$ D. $CH_3-\underset{\displaystyle CH_3}{\underset{|}{\overset{\displaystyle CH_3}{\overset{|}{C}}}}-$

35. Which of following alkenes has neither secondary nor tertiary hydrogens?

A) Iso - Butane C) n - Pentane

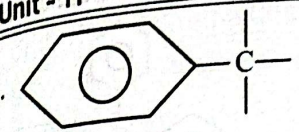
B) Pentane D) neo - pentane

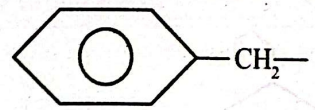
36. Chemically similar compounds differing by a - CH_2- are called

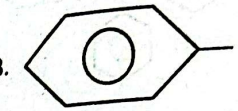
A) Isomers C) Isomor phous

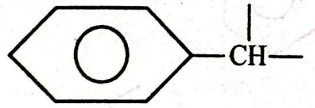
B) Homologues D) Isotopes

37. Which of the following represents the benzylene group?

A. 

C. 

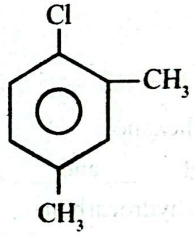
B. 

D. 

38. The reaction below is $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$

A) Addition reaction C) Cracking

B) Elimination reaction D) Substitution

39. The IUPAC name of 

A) 4 - Chloro - m- xylene

B) 4 - Chloro - 1,3 - dimethyl benzene

C) 1 - chloro - 2,4 - dimethyl benzene

D) 4 - chloro - 3 - methyltoluene.

40. All of the following are incorrect IUPAC names. **EXCEPT**

A) 2,2 - Dimethyl - 5 ethyl hexane

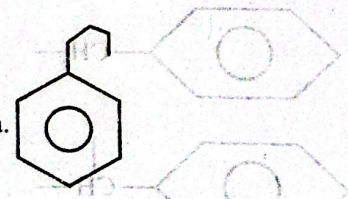
B) 3 - Ethyl - 4 , 4 - dimethyl hexane

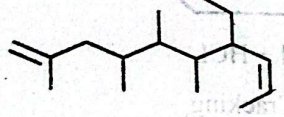
C) 4 - ethyl - 5,5 - dimethyl hexane

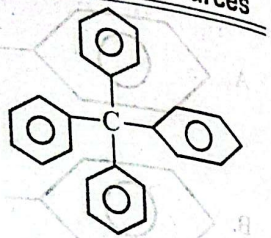
D) 2,2,3,3 - tetramethylpentane

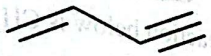
III. Short answer Items

41. Give the IUPAC name for

a. 

b. 

c. 

d. 

42. Draw the structural formula for
   a) Allyl benzene
   b) Hepta - 1,3,5 - triene
   c) 1,2,3 - Tricyano propane
43. Write the structure isomers of hexane
44. Hydrocarbons are composed of _____ and _____.
45. _____ is the simplest aromatic hydrocarbon

Answer

Part I

1. True	3. True	5. False	7. True	9. False
2. False	4. True	6. False	8. True	10. True

Part II

11. B	17. A	23. D	29. C	35. D
12. C	18. B	24. D	30. B	36. B
13. A	19. A	25. D	31. B	37. C
14. D	20. D	26. C	32. B	38. D
15. C	21. B	27. B	33. A	39. C
16. C	22. D	28. C	34. B	40. D

Part III

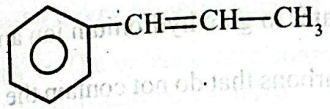
41. a) n - butyl benzene

   b) 2,4,5,6 - tetramethyl - 7- ethyl - 1,8 decadiene

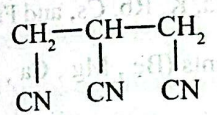
c) Terraphenyl methane

d) 3 - buten - 1 - yne

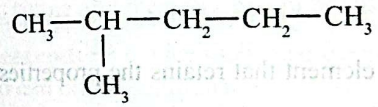
42.

a. 

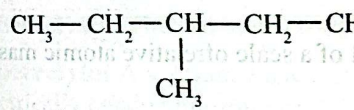
b. $CH_2=CH-CH=CH-CH=CH-CH_3$

c. 

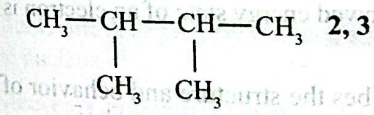
43. $CH_3-CH_2-CH_2-CH_2-CH_2-CH_3$ **Hexane**



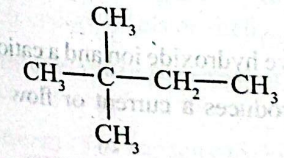
2 - methyl pentane



3 - methyl pentane



2, 3 - dimethyl butane



2, 2 - dimethyl butane

44. Hydrogen and carbon

45. Benzene

Accuracy: The agreement between the measured quantity and the accepted value.

Acid – base reaction: A reaction between an acid and a base.

Acid: A compound that ionizes in water to give hydronium ion and anion.

Aliphatic hydrocarbons: Hydrocarbons that do not contain the benzene group or the benzene ring.

Alkali metals: The Group 1A elements (Li, Na, K, Rb, Cs, and Fr).

Alkaline earth metals: The group 2A elements (Be , Mg , Ca , Sr , Ba and Ra).

Alpha particle: Positively charged particle ejected at high speeds from certain radioactive substances.

Atom: is the smallest particle of an element that retains the properties of that element.

Atomic Mass Unit (amu): The unit of a scale ofrelative atomic masses of the elements.

Atomic orbital: The wave for an allowed energy state of an electron is an atom.

Atomic theory: A theory that describes the structure and behavior of substances.

Base: A compound that ionizes in water to give hydroxide ion and a cation.

Battery: An electrochemical device that produces a current or flow of electrons at a constant voltage.

Beta: Particle electrons ejected at high speeds from certain radioactive nuclei.

Boiling point: The temperature at which the vapor pressure of a liquid equal to the external pressure on the liquid.

Bond energy: The energy that must be supplied to a gaseous molecule to separate two of its atoms.

Bond length: The distance between the centers of two bonded atoms in a molecule..

Bronsted - Lowry acid: A substance capable of donating a proton.

Bronsted-Lowry base: A substance capable of accepting a proton.

Chemistry: The science that studies the properties, compositions, and structures of substances, the transformations they undergo, and the energy that is released or absorbed during these processes

Colloid: A state intermediate between a solution and a suspension in which solute particles are large enough to scatter light but too small to settle out.

Coordinate covalent bond: A bond in which the pair of electrons is supplied by one of the two bonded atoms; also called a dative bond.

Double bond: A bond formed by sharing two pairs of electrons, one pair in a sigma and the other in a pi bonds

Electrode: A device such a same template or wire for conducting electrons into and out of solutions.

Electrolysis: The use of electrical energy to produce chemical change.

Electrolyte: A substance that ionizes in water or on melting to form an electrically conducting solution

Electron: A negatively charged subatomic particle found in space about the nucleus.

Electron configuration: Describes how electrons are distributed in their energy levels or shells

Endothermic process: A thermodynamics process in which heat flows into a system from its surroundings.

Energy: The capacity to do work and transfer of heat.

Enthalpy change (ΔH): Heat transferred at constant pressure

Exothermic process: A thermodynamic processing which heat flows from a system to its surroundings.

Extensive properties: Physical properties that depend on the amount of matter present.

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Heterogeneous mixture: The individual components of a mixture remain physically separated and can be seen as separate components.

Hypothesis: A tentative explanation of or production derived from experimental observations

Ideal solution: A solution that obeys Raoult's law.

Intensive properties: Physical properties that do not depend on the amount of matter present.

Intermolecular forces: Interactions between molecules, between or between molecules and ions.

Ion–dipole force: The attraction between an ion and a polar molecule.

Ion: An atom or group of atoms that has lost or gained one or more electrons.

Ionization energy: The minimum energy required to remove an electron from an isolated atom (or an ion) in its ground state.

Lewis acid: A substance that can accept a pair of electrons.

Lewis base: A substance that can donate a pair of electrons.

London forces: Very weak intermolecular forces occurring between two induced dipoles.

Melting point: The temperature at which the crystal lattice structure of a solid collapses and solid is converted to liquid.

Mixture: A contribution of two or more substances in which each substance retains its identity.

Molality (m): The number of moles of solute per kilogram of solvent.

Molar solubility: The number of moles of solute in one liter of a saturated solution (mol/L).

Molarity(M): The number of moles of solute per liter of solution.

Mole fraction(X): The ratio of the moles of a certain substance to the total moles in a mixture of substances.

Net ionic equation: A chemical equation in which only those ions undergoing chemical changes in the course of their action are represented.

Neutralization: An acid–base reaction that produces a neutral solution of salt and water.

Neutron: An electrically neutral subatomic particle found in the nucleus.

Non-electrolyte: A substance dissolves in water to form an electrically non – conducting solution.

Non - polar molecules: A molecule that does not possess a dipole moment.

Nucleus: The core of an atom made up of protons and neutrons.

Octet rule: An atom other than hydrogen tends to form bonds until it is surrounded by eight valance electron.

Polar covalent bond: In such a bond, the electrons spend more time in the vicinity of one atom than the other.

Polar molecule: A molecule for measuring the rotation of polarized light by optical isomers.

Precision: The agreement of repeated measurements of a quantity with one another

Proton: Positively charged subatomic particle found in the nucleus.

Salt: Anionic compound whose cation comes from a base and whose an ion comes from an acid.

Saturated solution: A stable solution in which the maximum amount of solute has been dissolved.

Significant figures: The number of meaningful digits in a measured or calculated quantity.

Single bond: Two atoms are held together by one electron pair.

Solute: The material dissolved in a solution.

Solution: A homogeneous mixture of two or more substances in a single phase.

Solvent: The medium in which a solute is dissolved to form a solution.

Temperature: A physical property that determine the direction of heat flow in an object on contact with another object.

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Theory: A unifying principle that explains a body off acts and the law based on them.

Triple bond: A bond formed by sharing three pairs of electrons, one is a sigma and the other two in pi bonds.

Unsaturated solution: A solution in which the concentration of dissolved solute is less than the saturation amount.

Valence shell electron: Pair repulsion (VSEPR) theory a model of structure about an atom in a molecule in which structural electron pairs are arranged so as to maximize the angles between them.

Vander Walls forces: Weak intermolecular forces not involving ions.

Volatility: The tendency of the molecules of a substance to escape from the liquid phase and enter the gas phase.

Weak acids: Acids that ionize only to a limited extent in water.

Weak bases: Bases that ionize only to a limited extent in water.

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